

# Chemistry Revision Themes

# Part I

## Themes:

# Atomic Structure

# Chemical Bonding

## Formulae, Equations & Moles

# The Periodic Table

## Oxidation & Reduction

# Inorganic Chemistry

																4.00 He 2					
6.94 Li 3		9.01 Be 4														10.81 B 5	12.01 C 6	14.01 N 7	16.00 O 8	19.00 F 9	20.18 Ne 10
22.99 Na 11		24.31 Mg 12														26.98 Al 13	28.09 Si 14	30.97 P 15	32.07 S 16	35.45 Cl 17	39.95 Ar 18
39.10 K 19	40.08 Ca 20	44.96 Sc 21	47.87 Ti 22	50.94 V 23	52.00 Cr 24	54.94 Mn 25	55.85 Fe 26	58.93 Co 27	58.69 Ni 28	63.55 Cu 29	65.39 Zn 30	69.72 Ga 31	72.64 Ge 32	74.92 As 33	78.96 Se 34	79.90 Br 35	83.80 Kr 36				
85.47 Rb 37	87.62 Sr 38	88.91 Y 39	91.22 Zr 40	92.91 Nb 41	95.94 Mo 42	99 Tc 43	101.07 Ru 44	102.91 Rh 45	106.42 Pd 46	107.87 Ag 47	112.41 Cd 48	114.82 In 49	118.70 Sn 50	121.75 Sb 51	127.60 Te 52	126.90 I 53	131.29 Xe 54				
132.91 Cs 55	137.33 Ba 56	138.91 La 57 *	178.49 Hf 72	180.95 Ta 73	183.84 W 74	186.21 Re 75	190.23 Os 76	192.22 Ir 77	195.09 Pt 78	196.97 Au 79	200.59 Hg 80	204.38 Tl 81	207.2 Pb 82	208.98 Bi 83	209 Po 84	210 At 85	222 Rn 86				
223 Fr 87	226.03 Ra 88	227.03 Ac 89 **	261 Rf 104	262 Db 105	266 Sg 106	264 Bh 107	277 Hs 108	268 Mt 109	271 Ds 110	272 Rg 111											

## CONTENT

First Year (AS) Chemistry guide: ■      Second Year (A2) Chemistry guide: ■  
 (NB: Consult *your* examining board syllabus to establish precise requirements)

### Part I:

#### **Theme One: Atomic Structure**

1

Components of the atom, Electron shells, Isotopes, Mass spectrometer,  
 The Bohr atom & electron transitions, Orbitals and relative orbital energies,  
 Electron configurations, Ionization energies, Electron affinities

#### **Theme Two: Chemical Bonding & Structures**

4

The noble gas structures, Ionic bonding, Properties of ionic substances,  
 Covalent Bonding, Properties of covalent substances, Pi & sigma bonds, Dative  
 Covalency, Polar molecules & hydrogen bonding, van der Waal forces,  
 Dual covalent-inorganic nature of some compounds, Metallic bonding,  
 Molecular geometry.

#### **Theme Three: Formulae, Equations & Moles**

12

Element symbols, Chemical formulae, Balanced chemical equations, The mole  
 Molar volume, Empirical & Molecular formulae, Percentage composition, Structural  
 formulae, Water of crystallisation, Reactions involving gases

#### **Theme Four: The Periodic Table**

19

Dimitri Mendeleev, Groups and Periods, Periodicity, Electrical conductivity, Melting  
 points & Boiling points, Ionisation energies, Electronic structures, Atom sizes, Isoelec-  
 tronic ions.

#### **Theme Five: Oxidation & Reduction**

23

Discovery of oxygen, Oxidation & reduction, Oxidation number, Blast furnace, Cata-  
 lytic converter, Redox, Electron transfer, Half equation, Application of oxidation  
 number, Balancing chemical equations.

#### **Theme Six: Inorganic Chemistry**

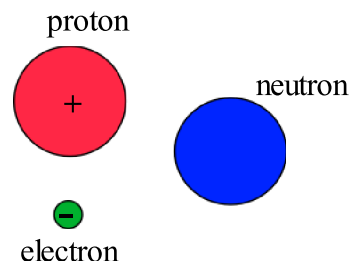
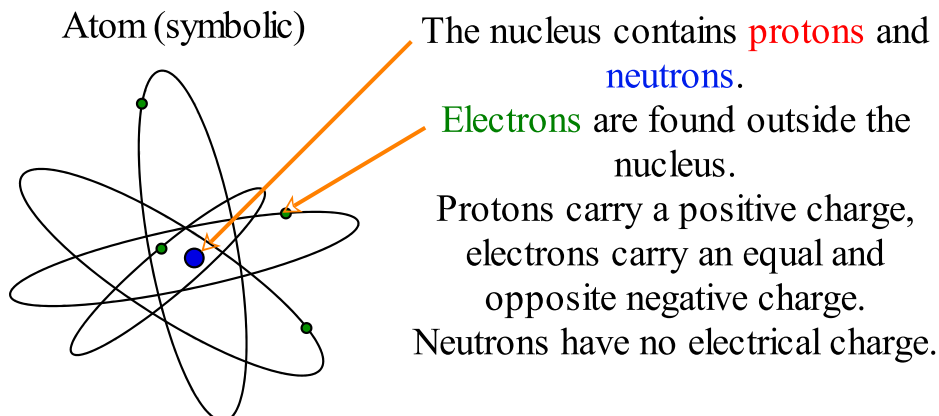
28

Halogens, The s-Block Elements, The Period 3 Elements, Transition Metals

# Theme One

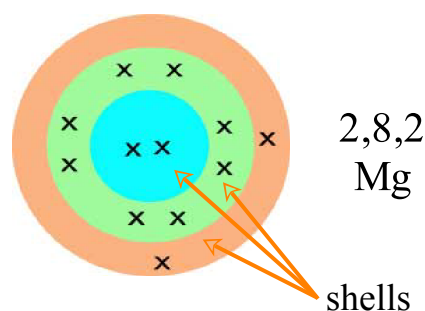
## Atomic Structure

The British physicist, E. Rutherford proposed the structure of the atom depicted symbolically below (1911).



Electrons are less than one two thousandth the mass of a proton. They are arranged in shells (principal energy levels) about the nucleus of the atom.

eg,



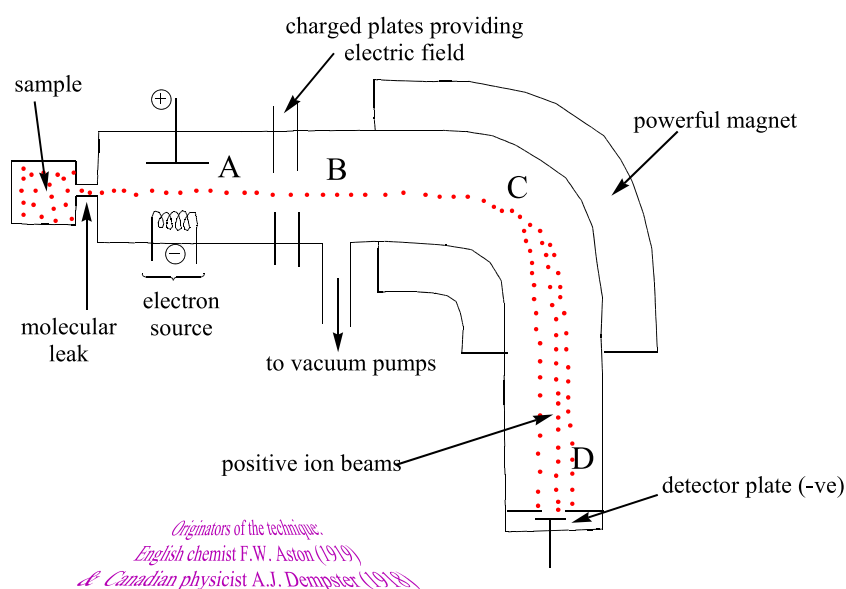
The English physicist, J.J. Thomson discovered the electron (1906)

**Isotopes:** atoms of the same element. They have the same number of protons but different numbers of neutrons & hence the same atomic numbers but different mass numbers.

eg,  $^{20}\text{Ne}$

$^{21}\text{Ne}$   $^{22}\text{Ne}$

neon isotopes (different mass numbers)



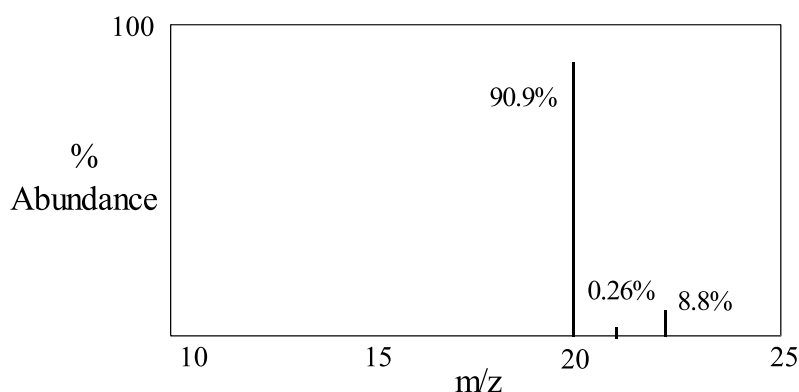
The **mass spectrometer** is used to separate and detect atoms, or molecules, of different mass. You need to know the basics of how it works and what it is used for.

[Refer to the bottom of the next page]

Referring to the letters A, B, C, D, state the positions, in the spectrometer, where the sample particles are, (i) ionised (ii) detected (iii) deflected (iv). accelerated?

# Atomic Structure

Mass Spectrum of Neon



Calculate the relative atomic mass of neon

Calculation:

$$(0.909 \times 20) + (0.0026 \times 21) + (0.088 \times 22) = 20.17$$

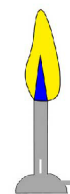
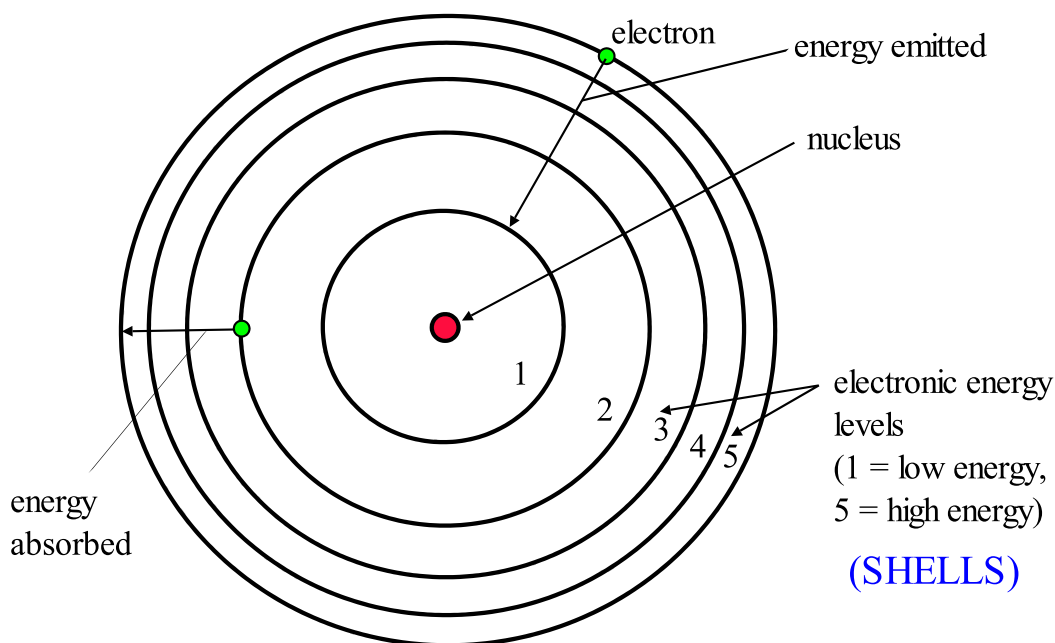
Neon discharge tubes were produced by the French chemist, Georges Claude in about 1910. Neon advertising signs became popular in the US & Europe in the 1920's.

## Atom model

The Danish physicist *Niels Bohr* proposed this model (1920)



A neon sign!



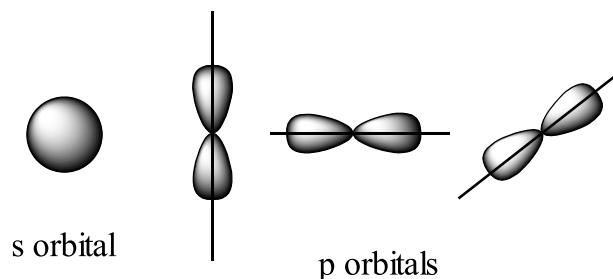
Why is it that sodium compounds give a bright yellow-orange colour when heated in a hot bunsen flame?

[Refer to the bottom of the next page for an answer.]

The shells are divided into subsidiary levels denoted s, p, d, f. The spaces occupied by the electrons in these subsidiary levels are called **orbitals**.

The orbitals have characteristic shapes

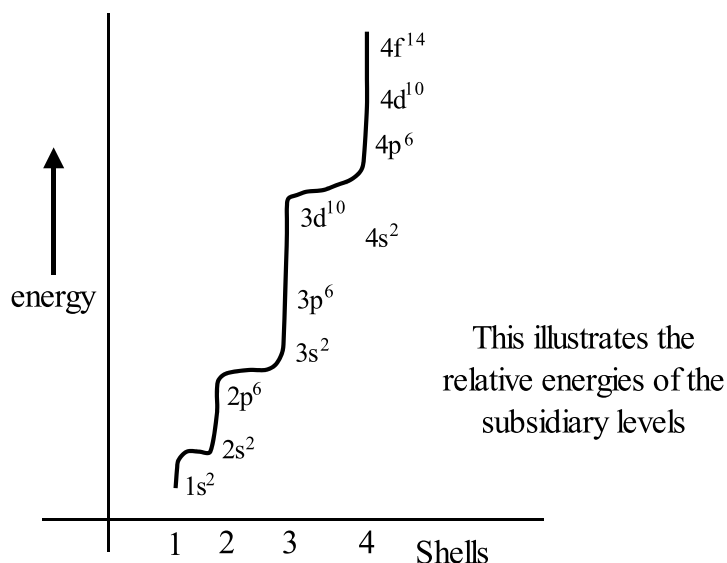
eg,



**[Mass Spectrometry:** The sample is vapourised at a suitable temperature and allowed to diffuse through a small hole known as the molecular leak into the ionisation chamber. Here the sample molecules pass through a beam of high energy electrons. On collision with the electrons the sample molecules lose an electron forming singly positively charged particles (some molecules are fragmented to yield additional, singly, positively charged particles). These charged particles are accelerated by electrically charged plates into a powerful magnetic field where they are separated according to mass. On reaching the detector plate the abundance of particles of a particular mass is recorded and the results displayed on chart paper as shown in the neon example above.]



# Atomic Structure



## Shorthand Electron Configurations:

**Carbon:** This has a total of six electrons arranged:  $1s^2, 2s^2, 2p^2$   
 Orbital boxes  $\longrightarrow$   $\uparrow\downarrow$   $\uparrow\downarrow$   $\uparrow\downarrow$   
**Sodium:** This has 11 electrons arranged:  $1s^2, 2s^2, 2p^6, 3s^1$   
**Iron:** This has 26 electrons arranged:  $[\text{Ar}], 3d^6, 4s^2$  where  $[\text{Ar}]$  is the electronic structure of argon ( $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$ )

**Electronic structure determines the chemical properties of an element.**

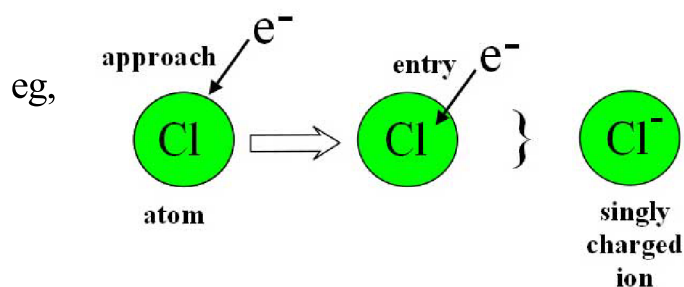
## Ionization Energy (Ionization Enthalpy)

The first ionization energy is defined as, *the energy required to remove one electron from each atom in one mole of gaseous atoms of an element.*

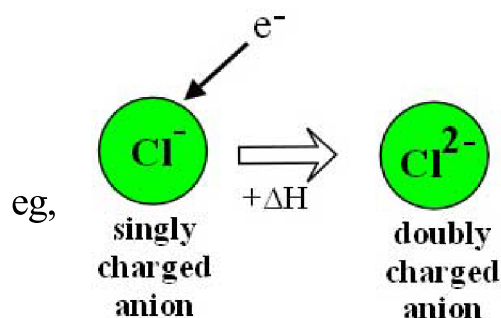
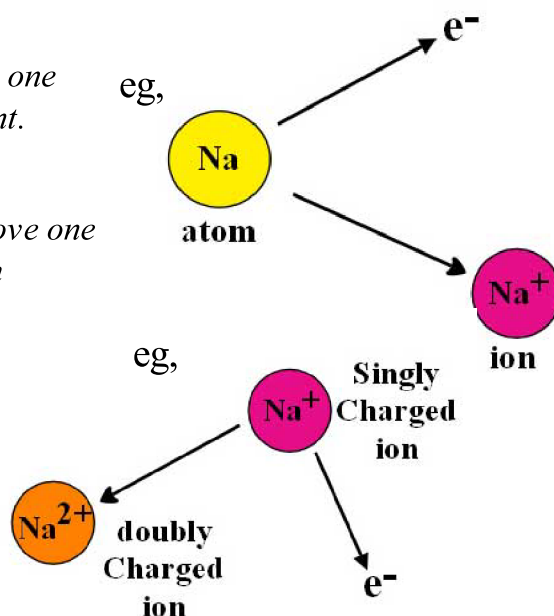
The second ionization energy is, *the minimum energy required to remove one electron from each unipositive ion in one mole of gaseous ions of an element.*

## Electron Affinity

The first electron affinity of an element is defined as, *the enthalpy change during the addition of one mole of electrons to one mole of isolated atoms of the element in the gaseous state.*



The second electron affinity is defined as, *the enthalpy change when one mole of electrons is added to one mole of isolated uninegative ions in the gaseous state. It is endothermic*



[Answer to question on previous page: Sodium Flame: Heat energy from the bunsen flame promotes an electron, in the sodium atom, to a higher energy level. Subsequently the electron returns to its original energy level and emits energy corresponding to yellow light in the visible region of the electromagnetic spectrum.]

# Theme Two

## Chemical Bonding & Structures

As a general rule:

**Atoms attain the electronic structures of the noble gases on chemical combination.**

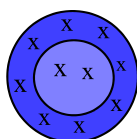
**This simple rule helps you to decide the type of bond formed between atoms on chemical combination and how many bonds will be produced.**

The German chemist, R. Abegg (1904) and later, the American chemists, G.N. Lewis & I. Langmuir (1916) introduced & developed the following descriptions of bonding.

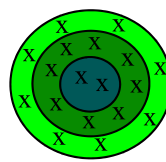
### Noble gases



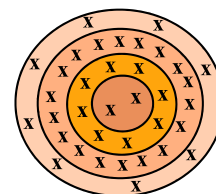
2  
He



2, 8  
Ne



2, 8, 8  
Ar

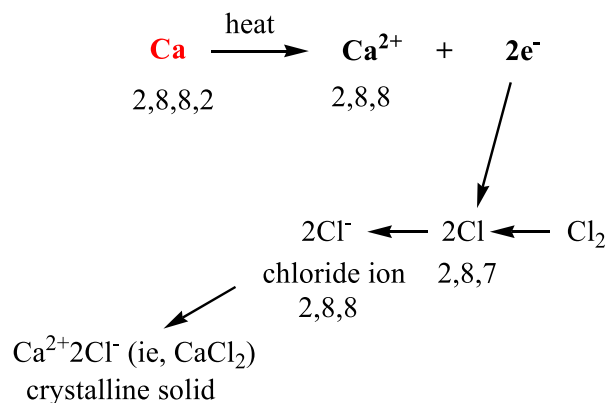
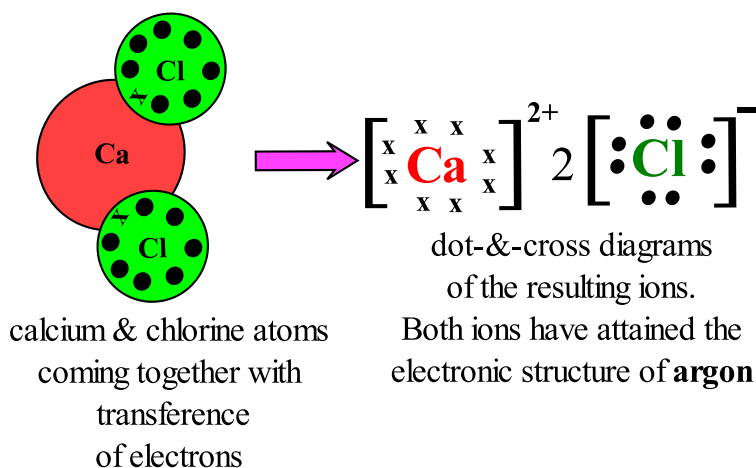


2, 8, 18, 8  
Kr

The **atomic orbital approach** is normally used to show how atoms achieve the stable electronic structures of the noble gases:

eg,

Heat calcium metal in chlorine gas. The two react **vigorously** forming calcium chloride which is a white crystalline solid.



Calcium chloride provides an example of a substance in which the bonding is typically **ionic**.

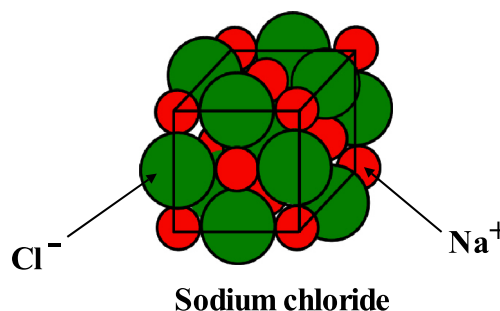
**The ionic bond involves strong electrostatic forces of attraction between the oppositely charge calcium and chloride ions.**

**Other ionic substances include: iron(II) sulphate ( $\text{Fe}^{2+}\text{SO}_4^{2-}$ ), potassium iodide ( $\text{K}^+\text{I}^-$ ), silver nitrate ( $\text{Ag}^+\text{NO}_3^-$ ) & magnesium carbonate ( $\text{Mg}^{2+}\text{CO}_3^{2-}$ ).**

## Chemical Bonding & Structures

In ionic substances the ions are arranged in regular geometric patterns. Many, for example, have a simple **cubic** structure.

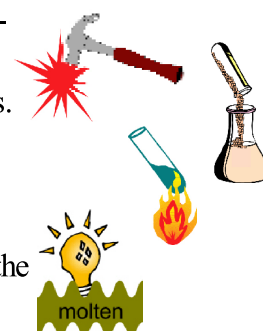
eg,



Looking at the bonding mechanism illustrated on the previous page it is clear that in order to form an ionic substance easily, the ionization energy of one element should be low compared with that of the other. In our example calcium has a **first ionization energy** of 590 kJ per mol whilst that of chlorine is 1260 kJ per mol. Also, the **first electron affinity** of chlorine (-364 kJ per mol) indicates that the atom will readily accept an electron. It is not surprising therefore that calcium reacts readily with chlorine & the reaction involves electron transfer and the formation of an ionic substance.

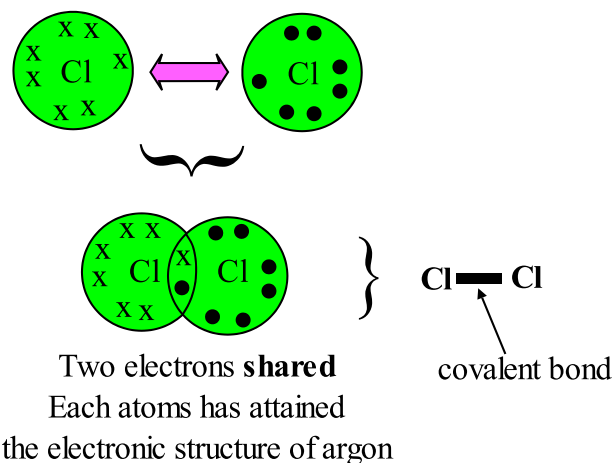
**Ionic substances are:-**

- Brittle crystalline solids.
- Dissolve in water.
- Melt at relatively high temperatures.
- Conduct electricity in the molten state.

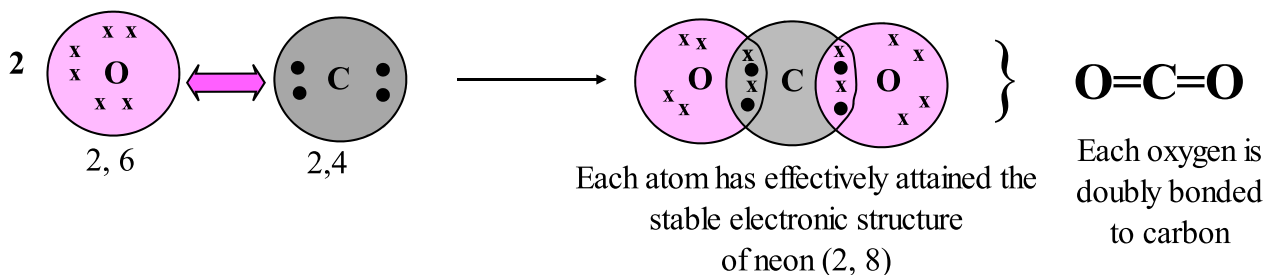
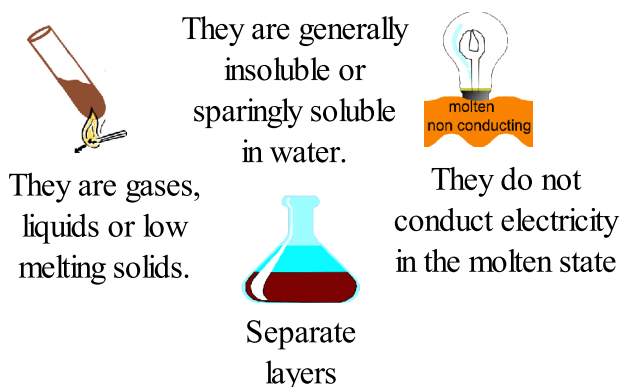


An alternative bonding mechanism results in the formation of **molecular substances**. eg. ammonia, hydrogen, oxygen, carbon dioxide, sulphur dioxide, water, ethanol, tetrachloromethane, trichloromethane, sulphur, phosphorus and urea.

**Covalent Bonding (dot-&-cross)**



**The covalent bond involves forces of attraction between the positively charged nuclei and the shared electrons.**

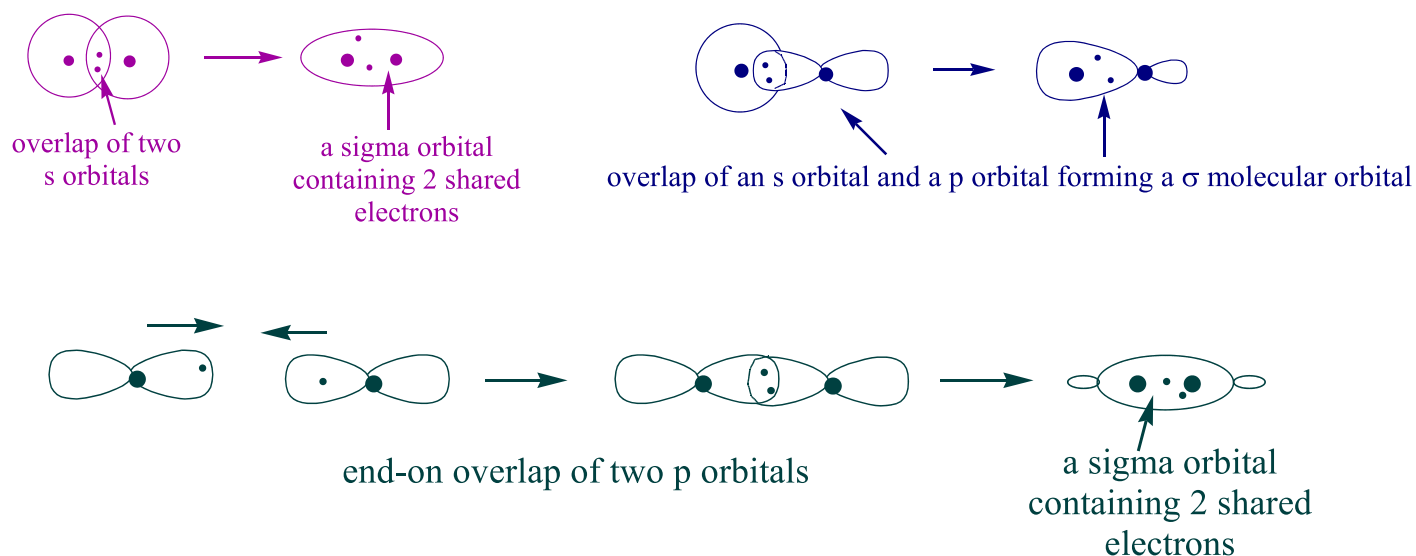


# Chemical Bonding & Structures

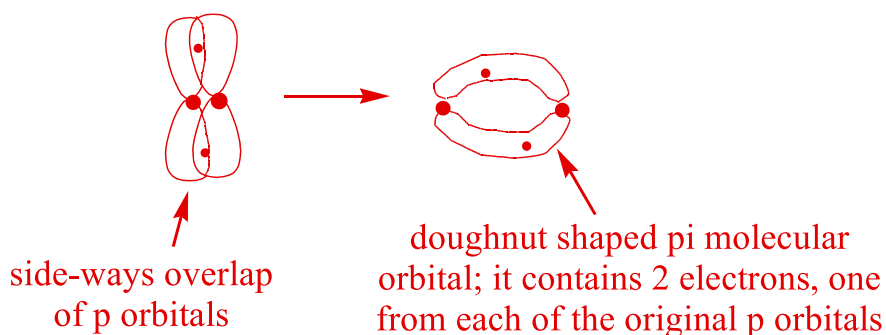
## Sigma ( $\sigma$ ) & Pi ( $\pi$ ) Bonds:

A sigma bond is an ordinary covalent bond. A pi bond is a weaker covalent bond

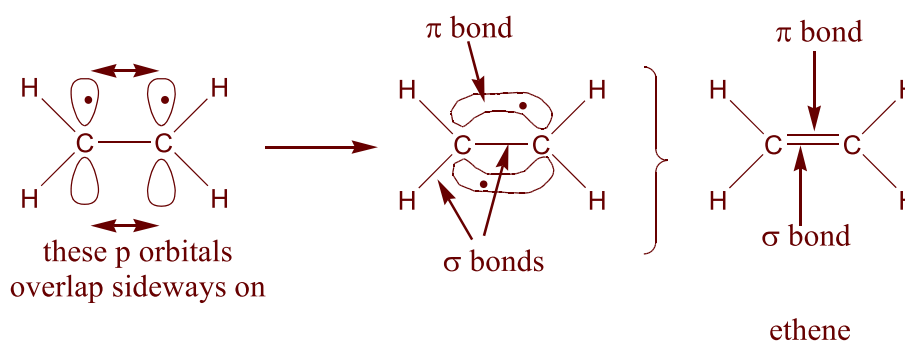
### Ways of forming sigma bonds:



### Pi bonds are formed by the side-on overlap of p orbitals:



The ethene molecule contains a sigma and a pi bond between the two carbon atoms:

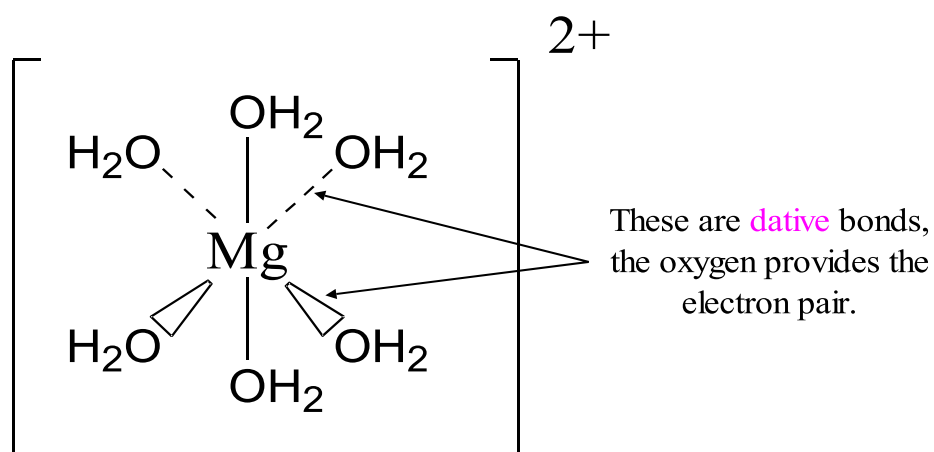
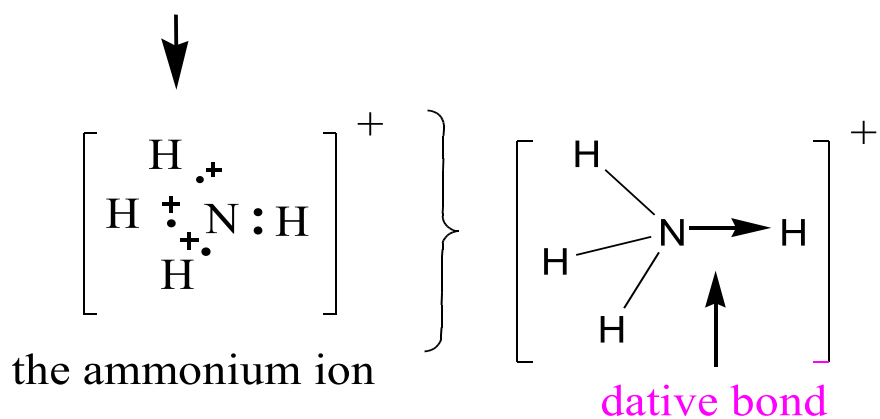
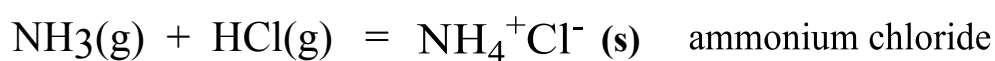


Ethene is a reactive gas. In its **addition** reactions it is the weak pi bond which is broken.

## Dative (or Co-ordinate ) Covalency

**In dative covalency both electrons in the bond are provided by just one of the bonded atoms.**

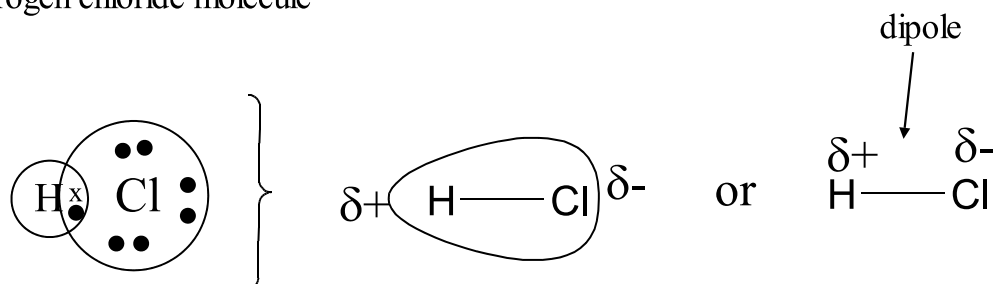
eg,



The hydrated magnesium ion

### The polar nature of some molecules.

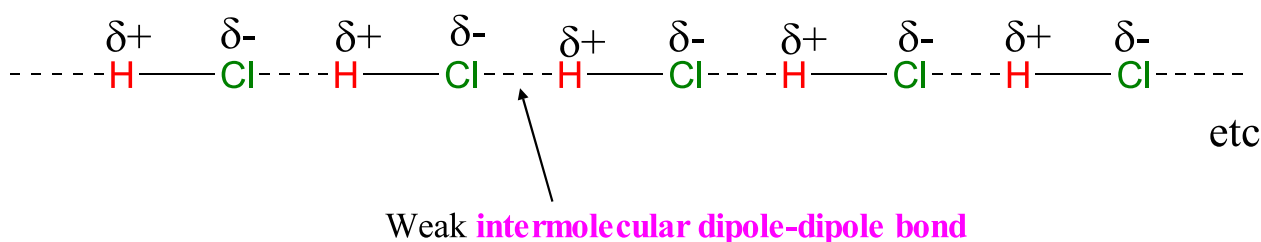
eg, hydrogen chloride molecule



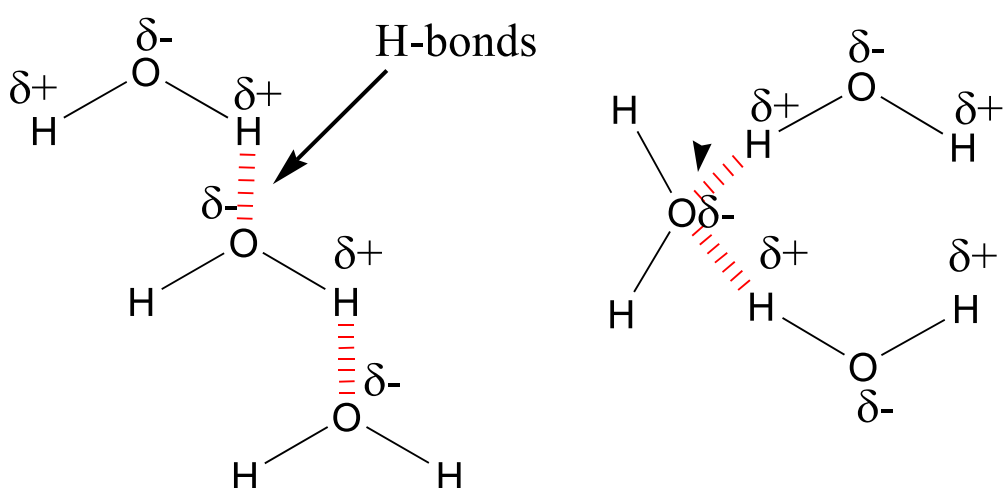
Chlorine has a greater affinity for electrons than hydrogen, we say it is more **electronegative** than hydrogen.

As a result of this, the hydrogen chloride molecule is **polarised** and has a **permanent electric dipole**.

In the liquid state the molecules join together forming long chains.



Another example: water



Hydrogen bonding in water

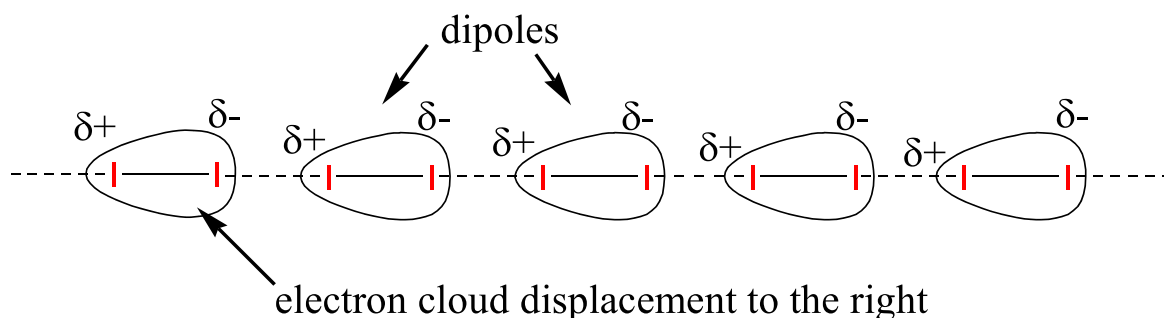
**Hydrogen bonding is responsible for the crystalline nature of ice and the relatively high boiling point of liquid water.**

(T.S. Moore & T.F. Winmill (1912) are credited with introducing the idea of hydrogen bonding)

## Chemical Bonding & Structures

Very weak temporary dipole-dipole forces can act between simple diatomic molecules. These forces are frequently known as **van der Waals forces** after the Dutch physicist J.D. van der Waals (1837-1923).

eg, Iodine,  $I_2$

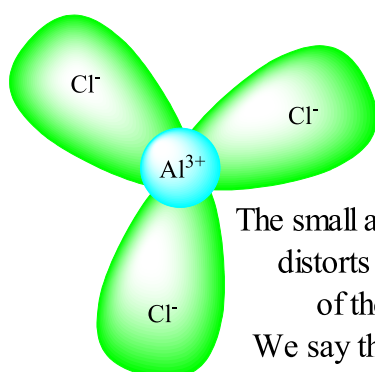


These weak dipole-dipole forces (van der Waals) account for the solid, crystalline, nature of iodine. van der Waals forces contribute to the physical state of all substances. Generally, the more electrons in atoms or molecules the stronger the van der Waals forces and the higher the melting and boiling points.

### The covalent nature of some inorganic substances.

Aluminium is a reactive metal. When heated, it combines with chlorine giving  $AlCl_3$ . We would expect this salt to have properties typical of ionic substances, but it does not! The small highly charged aluminium ion shares electrons with the chloride ions and this confers a good deal of covalent character on the salt.

Symbolic representation



The small aluminium ion (45 pm) distorts the electron cloud of the chloride ions. We say that the aluminium ion **polarises** the chloride ions

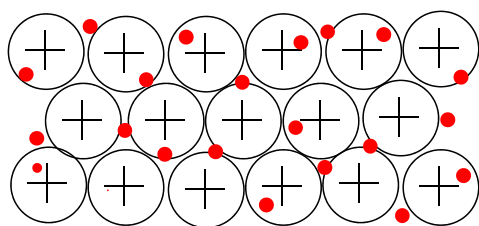
The Polish-born American chemist Kasimir Fajans is well known for his work relating to the covalent properties of compounds such as aluminium chloride, beryllium chloride and sodium iodide.

Aluminium chloride is a pale yellow solid with a relatively low melting point ( $180^\circ C$ ). It has a layer structure and forms dimers in the vapour state.

## Metallic Bonding:

Since metal atoms are spherical, in crystals of the metal they pack together like marbles in a box. The outer electrons of the atoms are loosely held and are free to move through the crystal lattice from atom to atom.

The metal atom nuclei are often described as being *embedded in a sea of electrons*.



looking down on the  
top layer of atoms

⊕ = nucleus of atom  
plus inner electrons  
  
• = outer, negatively  
charged, electron

This structure accounts for the characteristic properties of metals.

They readily conduct heat and electricity because the electrons carry negative charge and kinetic energy as they move freely between the relatively immobile nuclei and their inner electrons.

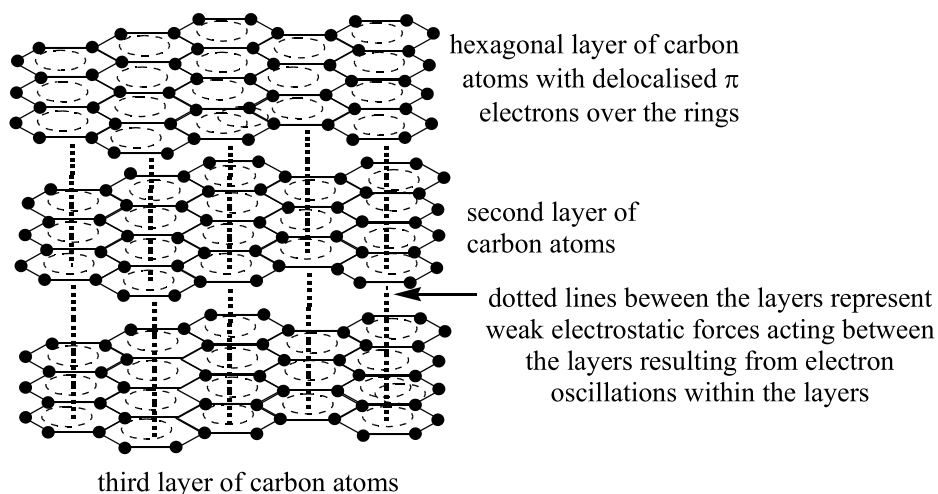
They are malleable and ductile because, under sufficient force, the atoms can slide over one another or squeeze in between one another and push one another aside so as to extend a row or layer of atoms.

Metals are generally, hard and high melting because the atoms are close packed and the bonding between the positively charged nuclei and the negatively charged electrons is effective and strong.

## Non-Metals:

Generally poor conductors of heat & electricity. This is because the outer electrons of the atoms are pretty immobile!

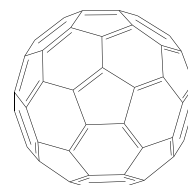
### An exception is graphite:



third layer of carbon atoms

Another form of carbon known  
as buckminsterfullerene discovered  
in 1985 by H.Kroto consists of  
C<sub>60</sub> particles.

The structure is unsaturated and can be hydrogenated.  
When combined with certain metals such as rubidium  
it becomes a good conductor of electricity.



C<sub>60</sub>  
(a 'buckyball')

Graphite has a soapy, slippery, feel. It is flaky and easily powdered.

Graphite will also conduct an electric current because there are unbonded electrons in the layers which carry the current.

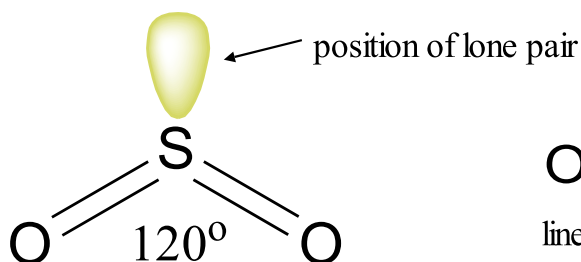


## Molecular Structures

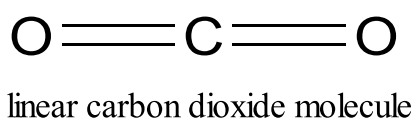
In these structures the bonding pairs and lone pairs are positioned as far apart as possible to reduce repulsive forces and give the molecules their most stable structures. This results in many of the molecules having three dimensional shapes.

**Order of repulsive forces between electron pairs:**

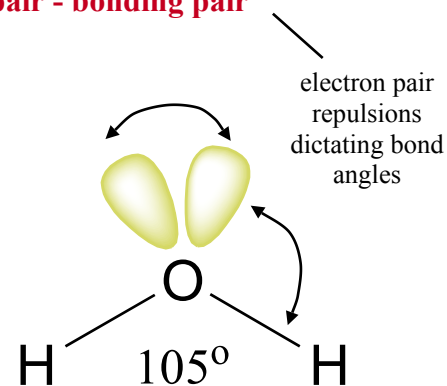
**lone pair - lone pair > lone pair - bonding pair > bonding pair - bonding pair**



V-shaped sulphur dioxide molecule

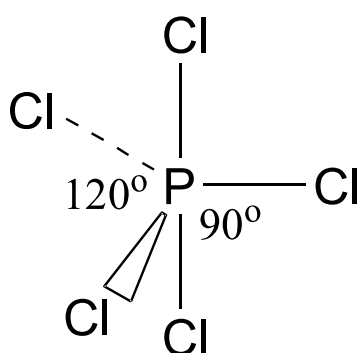


linear carbon dioxide molecule

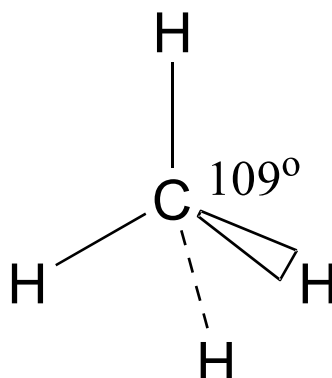


Bent water molecule

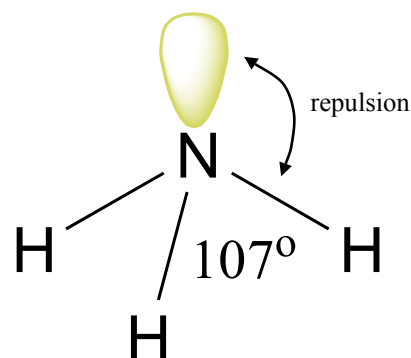
bond angle less than  $109^\circ$  due to the repulsive forces of the two lone pairs acting on one another and on the bonding pairs



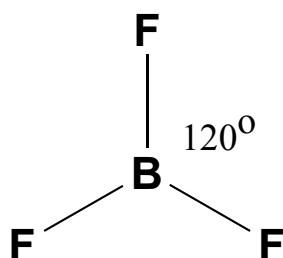
trigonal bipyramidal  
phosphorus pentachloride molecule



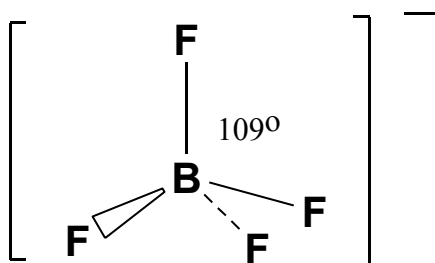
tetrahedral methane molecule



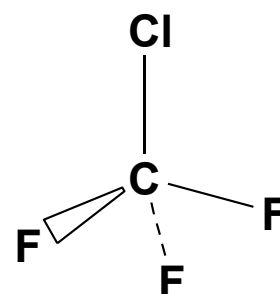
pyramidal ammonia molecule  
Bond angle a little less than  $109^\circ$  due to the influence of the single lone pair on the bonding pairs.



trigonal planar  
boron trifluoride



tetrahedral  
boron tetrafluoride ion  
(the 4th F bonds as a F<sup>-</sup> donating a pair of electrons into an empty 2p orbital of the boron)



tetrahedral  
chlorotrifluoromethane  
(Cl-C-F angle slightly larger than  $109^\circ$  due to larger size of Cl atom compared with F atom)

# Theme Three

## Formulae, Equations & Moles

### Formulae & Equations - the chemists shorthand

The elements have unique names and symbols.

eg,

Element Name	Chemical Symbol	Element Name	Chemical Symbol
Aluminium	Al	Neon	Ne
Argon	Ar	Nickel	Ni
Bromine	Br	Nitrogen	N
Calcium	Ca	Oxygen	O
Carbon	C	Phosphorus	P
Chlorine	Cl	Platinum	Pt
Copper	Cu	Potassium	K
Fluorine	F	Silicon	Si
Gold	Au	Silver	Ag
Hydrogen	H	Sulphur	S
Iodine	I	Tin	Sn
Iron	Fe	Titanium	Ti
Lead	Pb	Tungsten	W
Lithium	Li	Uranium	U
Magnesium	Mg	Vanadium	V
Manganese	Mn	Xenon	Xe
Mercury	Hg	Zinc	Zn

Often, the symbol is the first one or two letters in the English name, however, sometimes letters from the Latin names are used, eg, iron (Fe, Latin, ferrum), copper (Cu, Latin, cuprum), sodium (Na, Latin, natrium), lead (Pb, Latin, plumbum). The symbol K, for potassium, derives from the Arabic, *kali*, meaning potash. Notice that all of the symbols begin with a capital letter; if there is a second letter it is lower case (eg, Cl not CL).

**A compound is defined as, a substance which contains two or more elements chemically combined together.**

**The way in which the elements combine is dealt with in the theme, *chemical bonding*.**

**The *formula* of the compound indicates the proportions in which the elements are combined,**

**eg,**

Calcium oxide: CaO: Calcium and oxygen atoms combined in a 1:1 ratio.

Sodium oxide: Na<sub>2</sub>O: Sodium and oxygen atoms combined in a 2:1 ratio.

Methane: CH<sub>4</sub>: Carbon and hydrogen atoms combined in a 1:4 ratio.

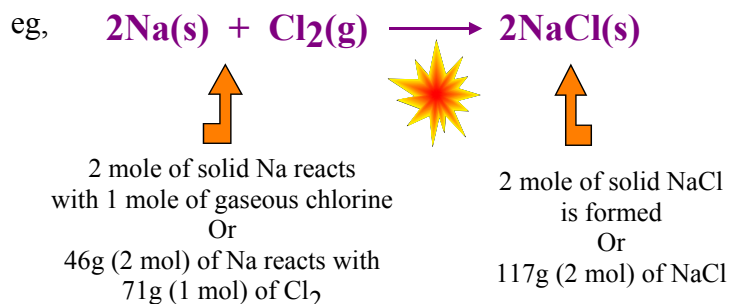
Sulphuric acid: H<sub>2</sub>SO<sub>4</sub>: Hydrogen, sulphur and oxygen atoms combined in a 2:1:4 ratio.

Iron chloride: FeCl<sub>2</sub>: Iron and chlorine atoms combined in a 1:2 ratio.

## Formulae, Equations & Moles

### Balanced chemical equations indicate:

- How substances can be made.
- Reacting mole ratios.
- Physical states of reactants and products.

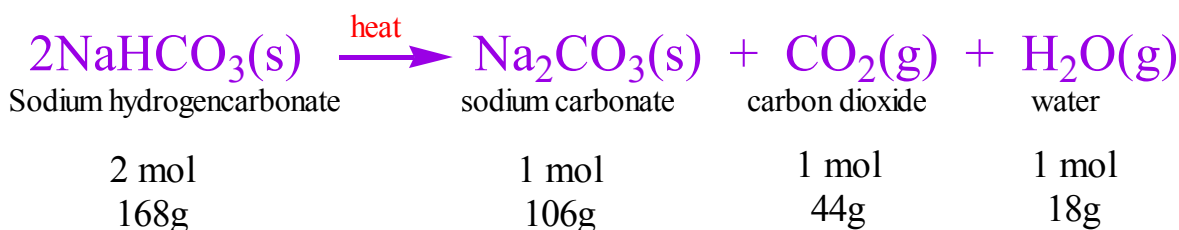


A reaction like this one between sodium and chlorine goes very readily and the yield of sodium chloride is very good (ie, all of the reactants are converted to product, 100% yield). However if, in an experiment, only 100 g of sodium chloride was isolated (from the given amounts of reactants), the % yield would be,  $(100 / 117) \times 100 = 85.47\%$ .

By the way!

A **mole** (Latin for *heap*) of an element or compound or ion is that *number* of fundamental particles of the substance as there are atoms in 12 g of the carbon-12 isotope,  
ie,  $6.02 \times 10^{23} \text{ mol}^{-1}$  (Avogadro's number)

Also, the molar mass of an element or compound provides one mole.



Sodium hydrogencarbonate (sodium bicarbonate) is a constituent of **baking powder**. It is used as a *raising agent* since it readily decomposes on gentle heating to provide carbon dioxide.

One mole of the hydrogencarbonate can provide 44 g (or **24 dm<sup>3</sup> measured at rtp**) of CO<sub>2</sub>.

The residual anhydrous sodium carbonate is stable to further heating.



Remember:

**1 mole of any gas occupies 24 dm<sup>3</sup> at room temperature & pressure (20°C, 1 atm).**

(Note: This can be confirmed by substituting,  $P = 1 \text{ atm}$ ,  $n = 1 \text{ mol}$ ,  $R = 0.0821 \text{ dm}^3 \text{ atm mol}^{-1} \text{ K}^{-1}$ ,  $T = 293 \text{ K}$  into the *ideal gas equation* and calculating  $V \text{ (dm}^3\text{)}$ .  
Ideal gas equation given on page 18.)

**Note:** This decomposition reaction is the final stage in the **Solvay process** for manufacturing sodium carbonate. On the face of it, it seems rather uneconomic since some of the carbon is lost as carbon dioxide instead of forming carbonate. The so called **% atom economy**, in terms of sodium carbonate, would be only,  $(106 / 168) \times 100 = 63.1\%$ . However, the carbon dioxide is not wasted since it is recycled and used in the *first* stage of the Solvay process!

## Formulae, Equations & Moles

### Empirical & Molecular Formulae

**Empirical formula (EF)** - the simplest formula for a substance- states only the **type** and **ratio** of atoms.

eg,

The hydrocarbon gas **ethyne (acetylene)** EF = **CH**

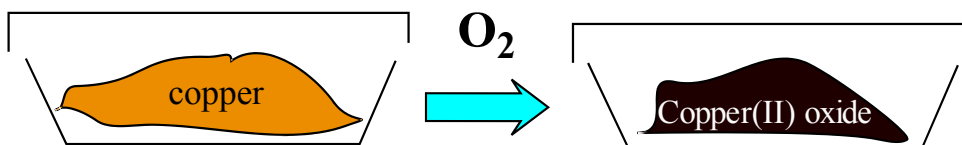
ie, comprises carbon and hydrogen in the atom ratio 1:1.

However, the formula of one molecule of the substance (the **molecular formula**) is



Mass spectrometry may be used to establish empirical and molecular formulae although other methods of analysis are also available.

### Analysis of copper(II) oxide



Mass of crucible & lid	20.25g
Mass of crucible, lid & copper	20.53g
Mass of crucible, lid & copper oxide	20.60g

Element	Mass	Mole
copper	0.28g	$0.28/63.5 = 0.0044 \text{ mol}$
oxygen	0.07g	$0.07/16 = 0.0044 \text{ mol}$

Copper & oxygen combine in equal amounts, ie, 1:1 mole ratio. Empirical formula of the oxide is CuO.

The percentage of copper in the oxide is  $(0.28/0.35) \times 100 = 80\%$ . Percentage oxygen is  $20\%$  ( $(0.07/0.35) \times 100 = 20\%$ ).

## Formulae, Equations & Moles

### Percentage Composition & Empirical Formulae

eg,

Hydrogen Peroxide	Percentage(%)
Hydrogen	5.88
Oxygen	94.12

100g of hydrogen peroxide is comprised of 5.88g hydrogen combined with 94.12g of oxygen

To find the mole amounts divide the masses by the relative atomic masses

Hydrogen Peroxide	Percentage(%)	Mole
Hydrogen	5.88	$5.88/1 = 5.88$
Oxygen	94.12	$94.12/16 = 5.88$

Since mole ratios are directly proportional to atom ratios the ratio of hydrogen atoms to oxygen atoms in hydrogen peroxide is 1:1

Therefore EF = HO

Mass spectrometry shows that hydrogen peroxide has a relative molecular mass of 34.

This is twice the EF therefore the molecular formula of hydrogen peroxide is



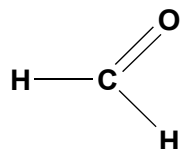
**A gaseous compound gave on analysis:**

Element	Carbon	Hydrogen	Oxygen
%	40.04	6.66	53.30
Divide by ram	$40.04/12$	$6.66/1$	$53.30/16$
Mole ratios	3.33	6.66	3.33
Atom ratios	$3.33/3.33 = 1$	$6.66/3.33 = 2$	$3.33/3.33 = 1$

Empirical formula = **CH<sub>2</sub>O**

Mass spectrometry shows that this compound has a relative molecular mass of 30, therefore its molecular formula is the same as its empirical formula.

It *structural formula* is:



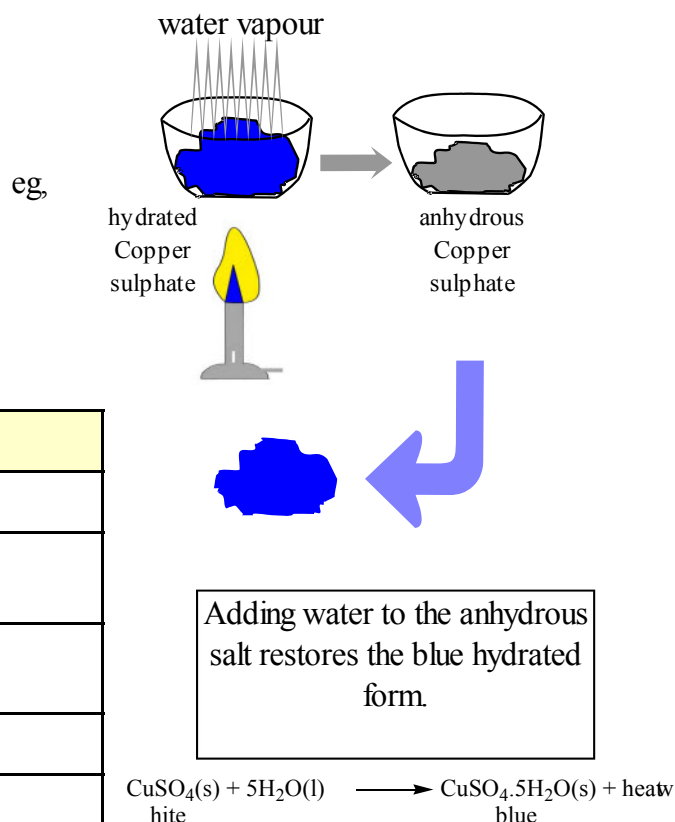
## Water of Crystallisation

Salts often have water incorporated in their crystal lattices.

eg  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  (hydrated copper sulphate)  
 $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$  (hydrated magnesium sulphate)  
 $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$  (hydrated sodium carbonate)

It is often possible to determine the amount of water of crystallisation by heating the salt in a crucible and recording the overall loss in mass.

Stage	Mass
Crucible + lid	19.84g
Crucible + lid + hydrated copper sulphate	25.84g
Crucible + lid + anhydrous copper sulphate	23.68g
Hydrated copper sulphate	$25.84 - 19.84 = 6.00\text{g}$
Mass of water lost	$25.84 - 23.68 = 2.16\text{g}$
% water of crystallisation	$(2.16/6.00) \times 100 = 36.00\%$



We can also use this data to establish the formula of the hydrated salt:

Anhydrous copper sulphate	3.84g
Water of crystallisation	2.16g
Mole of anhydrous copper sulphate	$3.84/159.5 = 0.024$
Mole of water of crystallisation	$2.16/18 = 0.12$
Mole ratio: copper sulphate:water	0.024:0.12 ie, 1:5

Therefore,  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

# Chemical reactions involving gases

eg.



Calcium carbonate

Carbon dioxide

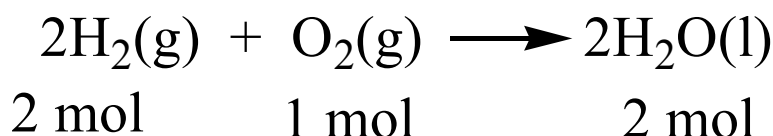
- main constituent of limestone

1 mol  $\text{CaCO}_3$  gives 1 mol  $\text{CO}_2$

At room temperature and pressure (20°C & 1 atm) this amount of gas occupies 24 dm<sup>3</sup>.

All gases at room temperature and pressure occupy 24 dm<sup>3</sup> (Avogadro's Law)

Amadeo Avogadro  
Italian physicist  
(1776 - 1856)



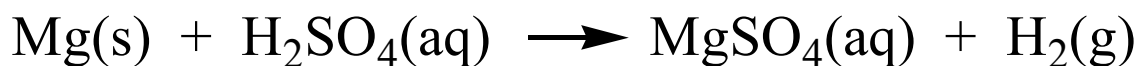
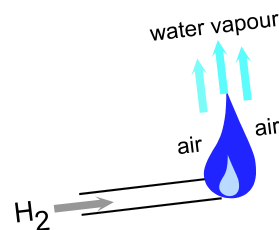
2 mol

1 mol

2 mol

Highly exothermic!

36g (approx. 36 cm<sup>3</sup>)



1 mol

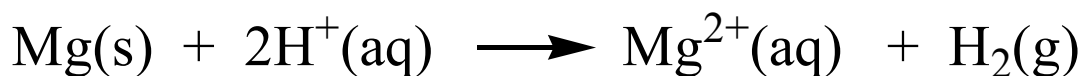
2g

24 dm<sup>3</sup>

Metals in the top half of the reactivity series react with dilute sulphuric and hydrochloric acids producing a salt and hydrogen.

The ionic form of the above equation shows that the sulphate ions are *spectator ions* (ie, do not participate chemically).

an ionic equation

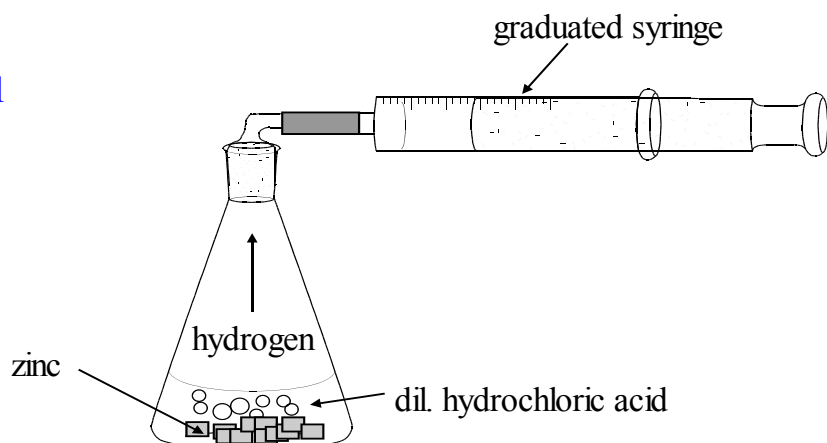


provided by the sulphuric acid

# Formulae, Equations & Moles

Consider these experiments:

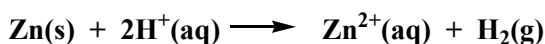
1



36.7 cm<sup>3</sup> of H<sub>2</sub> finally collected

What mass of Zn would have been required to produce this volume of hydrogen at room temperature and pressure?

**Ionic equation:**



1 mol

1 mol, 24 dm<sup>3</sup>

$$\text{Mol of hydrogen collected} = \frac{0.0367}{24} = 0.00153$$

Then, mol of zinc reacted = 0.00153

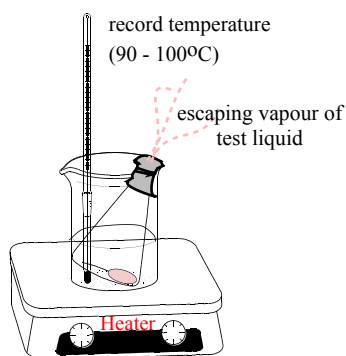
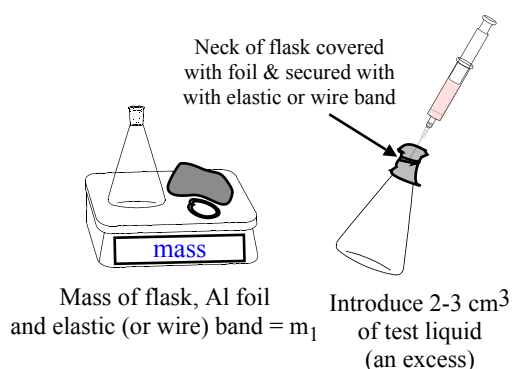
now, 1 mol Zn = 65.4 g

therefore, 0.00153 mol Zn = 0.00153 x 65.4 = 0.100 g

2

**Molar mass of a volatile liquid? (using the Dumas method & the ideal gas equation).**

**Jean Dumas (about 1820)**



Immerse flask in hot water and heat until the test liquid disappears (keep foil out of the water!)



Remove flask from hot water, allow to cool, dry flask and weigh, mass = m<sub>2</sub>

volume of flask = V m<sup>3</sup>  
(measured by filling to the top with water from a measuring cylinder)  
This equals the volume of the vapour in the flask

Using barometer, measure air pressure = P Nm<sup>-2</sup>

Ideal gas equation

$$\text{PV} = nRT$$

$$\text{PV} = \frac{(m_2 - m_1)}{M} RT$$

$$M = \frac{(m_2 - m_1) RT}{PV}$$

m = sample mass (g)

M = molar mass (g mol<sup>-1</sup>)

n = moles

In the experiment,

$$P = 101\,400 \text{ Nm}^{-2}$$

$$V = 1.40 \times 10^{-4} \text{ m}^3$$

$$R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$$

$$T = (90 + 273) \text{ K}$$

$$m_1 = 60.045 \text{ g}$$

$$m_2 = 60.460 \text{ g}$$

Molar mass (M) =

$$\frac{0.415 \times 8.314 \times 363}{101\,400 \times 1.40 \times 10^{-4}} = 88.2 \text{ g mol}^{-1}$$

(The test liquid was *hexane* (bp 69°C, M = 86 g mol<sup>-1</sup>).  
What is the % error in the result?)



# Theme Four

## The Periodic Table

### The Periodic Table & Periodicity

The Periodic table is a list of the known elements arranged in order of atomic number starting with hydrogen.


The first periodic table was published in 1869 by the Russian chemist, [Dimitri Mendeleev](#). The German chemist [Lothar Meyer](#) published similar result (independently) in 1870. The periodic table was a major innovation since it helped to make sense of the mass of chemical knowledge available at the time. A short time before Mendeleev's

death, in 1907, it is said that he missed receiving the Nobel Prize for chemistry by one vote! The prize went to the French chemist, Moissan, for his achievement in isolating fluorine.


Mendeleev's table listed the elements in order of atomic weight whereas today we use atomic number. In honour of Mendeleev, element 101 was named mendelevium (Md) when it was discovered in 1955.

### An example of the modern form of the periodic table:

																1.01 H 1												4.00 He 2	
6.94 Li 3		9.01 Be 4										10.81 B 5		12.01 C 6		14.01 N 7		16.00 O 8		19.00 F 9		20.18 Ne 10							
22.99 Na 11		24.31 Mg 12										26.98 Al 13		28.09 Si 14		30.97 P 15		32.07 S 16		35.45 Cl 17		39.95 Ar 18							
39.10 K 19		40.08 Ca 20		44.96 Sc 21	47.87 Ti 22	50.94 V 23	52.00 Cr 24	54.94 Mn 25	55.85 Fe 26	58.93 Co 27	58.69 Ni 28	63.55 Cu 29	65.39 Zn 30	69.72 Ga 31	72.64 Ge 32	74.92 As 33	78.96 Se 34	79.90 Br 35	83.80 Kr 36										
85.47 Rb 37		87.62 Sr 38		88.91 Y 39	91.22 Zr 40	92.91 Nb 41	95.94 Mo 42	99 Tc 43	101.07 Ru 44	102.91 Rh 45	106.42 Pd 46	107.87 Ag 47	112.41 Cd 48	114.82 In 49	118.70 Sn 50	121.75 Sb 51	127.60 Te 52	126.90 I 53	131.29 Xe 54										
132.91 Cs 55		137.33 Ba 56		138.91 La 57 *	178.49 Hf 72	180.95 Ta 73	183.84 W 74	186.21 Re 75	190.23 Os 76	192.22 Ir 77	195.09 Pt 78	196.97 Au 79	200.59 Hg 80	204.38 Tl 81	207.2 Pb 82	208.98 Bi 83	209 Po 84	210 At 85	222 Rn 86										
223 Fr 87		226.03 Ra 88		227.03 Ac 89 **	261 Rf 104	262 Db 105	266 Sg 106	264 Bh 107	277 Hs 108	268 Mt 109	271 Ds 110	272 Rg 111																	
				* Lanthanides																									
				140.12 Ce 58	140.91 Pr 59	144.24 Nd 60	147 Pm 61	150.36 Sm 62	151.69 Eu 63	157.25 Gd 64	158.93 Tb 65	165.50 Dy 66	164.93 Ho 67	167.26 Er 68	168.93 Tm 69	173.04 Yb 70	174.97 Lu 71												
				* * Actinides																									
				232.04 Th 90	231.04 Pa 91	238.03 U 92	237.05 Np 93	244 Pu 94	243 Am 95	247 Cm 96	247 Bk 97	251 Cf 98	252 Es 99	257 Fm 100	257 Md 101	259 No 102	261 Lr 103												

				
Alkali metals	Alkaline Earth metals	Transition elements	Halogens	Noble gases

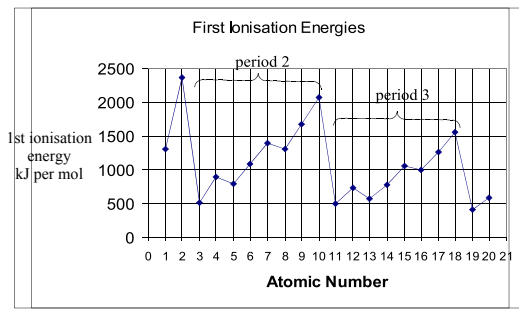
## The Periodic Table



The vertical columns are called **groups**. Elements in a group have similar chemical properties because they have similar outer electronic structures.

This group, or family, of elements is known as the **alkali metals**

**Periodicity** means *repeating pattern*. It refers to the way in which the properties of the elements change, in a predictable way, in going across the periods and down the groups.



As a general rule, the metallic elements have high electrical conductivity. The non metals have low electrical conductivity.

As a general rule, the 1st ionisation energies *increase* across a period. This matches a gradual *decrease* in atom size. The change in ionisation energy is not *perfectly* smooth due to predictable variations in electronic stability as the energy levels are filled.

# The Periodic Table

s-block										1.01 H 1	p-block										4.00 He 2
6.94 Li 3	9.01 Be 4	d-block										10.81 B 5	12.01 C 6	14.01 N 7	16.00 O 8	19.00 F 9	20.18 Ne 10				
22.99 Na 11	24.31 Mg 12											26.98 Al 13	28.09 Si 14	30.97 P 15	32.07 S 16	35.45 Cl 17	39.95 Ar 18				
39.10 K 19	40.08 Ca 20											44.96 Sc 21	47.87 Ti 22	50.94 V 23	52.00 Cr 24	54.94 Mn 25	55.85 Fe 26	58.93 Co 27	58.69 Ni 28	63.55 Cu 29	65.39 Zn 30
85.47 Rb 37	87.62 Sr 38	88.91 Y 39	91.22 Zr 40	92.91 Nb 41	95.94 Mo 42	99 Tc 43	101.07 Ru 44	102.91 Rh 45	106.42 Pd 46	107.87 Ag 47	112.41 Cd 48	114.82 In 49	118.70 Sn 50	121.75 Sb 51	127.60 Te 52	126.90 I 53	131.29 Xe 54				
132.91 Cs 55	137.33 Ba 56	138.91 La 57 *	178.49 Hf 72	180.95 Ta 73	183.84 W 74	186.21 Re 75	190.23 Os 76	192.22 Ir 77	195.09 Pt 78	196.97 Au 79	200.59 Hg 80	204.38 Tl 81	207.2 Pb 82	208.98 Bi 83	209 Po 84	210 At 85	222 Rn 86				
223 Fr 87	226.03 Ra 88	227.03 Ac 89 **	261 Rf 104	262 Db 105	266 Sg 106	264 Bh 107	277 Hs 108	268 Mt 109	271 Ds 110	272 Rg 111											
* Lanthanides		140.12 Ce 58	140.91 Pr 59	144.24 Nd 60	147 Pm 61	150.36 Sm 62	151.69 Eu 63	157.25 Gd 64	158.93 Tb 65	165.50 Dy 66	164.93 Ho 67	167.26 Er 68	168.93 Tm 69	173.04 Yb 70	174.97 Lu 71						
		** Actinides		232.04 Th 90	231.04 Pa 91	238.03 U 92	237.05 Np 93	244 Pu 94	243 Am 95	247 Cm 96	247 Bk 97	251 Cf 98	252 Es 99	257 Fm 100	257 Md 101	259 No 102	261 Lr 103				

# The Periodic Table

You will need to be able to write the electronic structures of these elements given their atomic numbers:

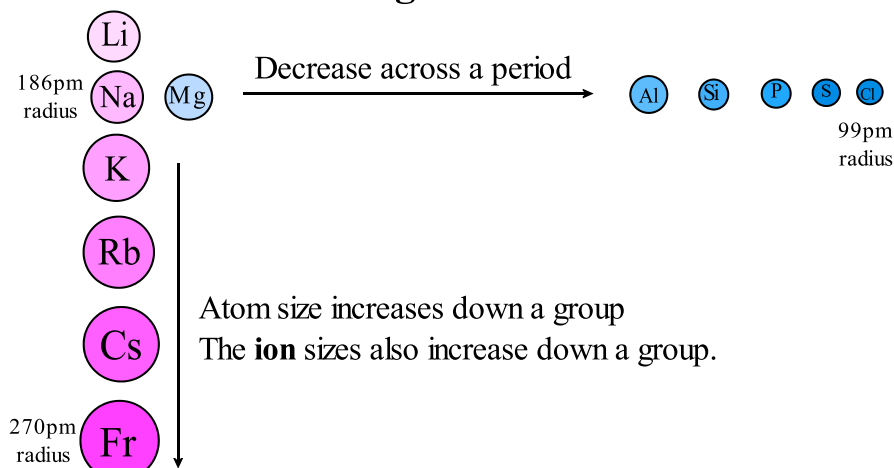
Element symbol	Atomic number	Electronic Structure							
		1s	2s	2p	3s	3p	3d	4s	4p
H	1	1							
He	2	2							
Li	3	2	1						
Be	4	2	2						
B	5	2	2	1					
C	6	2	2	2					
N	7	2	2	3					
O	8	2	2	4					
F	9	2	2	5					
Ne	10	2	2	6					
Na	11	2	2	6	1				
Mg	12	2	2	6	2				
Al	13	2	2	6	2	1			
Si	14	2	2	6	2	2			
P	15	2	2	6	2	3			
S	16	2	2	6	2	4			
Cl	17	2	2	6	2	5			
Ar	18	2	2	6	2	6			
K	19	2	2	6	2	6		1	
Ca	20	2	2	6	2	6		2	
Sc	21	2	2	6	2	6	1	2	
Ti	22	2	2	6	2	6	2	2	
V	23	2	2	6	2	6	3	2	
Cr	24	2	2	6	2	6	5	1	
Mn	25	2	2	6	2	6	5	2	
Fe	26	2	2	6	2	6	6	2	
Co	27	2	2	6	2	6	7	2	
Ni	28	2	2	6	2	6	8	2	
Cu	29	2	2	6	2	6	10	1	
Zn	30	2	2	6	2	6	10	1	
Ga	31	2	2	6	2	6	10	2	1
Ge	32	2	2	6	2	6	10	2	2
As	33	2	2	6	2	6	10	2	3
Se	34	2	2	6	2	6	10	2	4
Br	35	2	2	6	2	6	10	2	5
Kr	36	2	2	6	2	6	10	2	6

Examine Cr and Cu electronic structures.

Do they buck the trend???



## Changes in atom size:



## Isoelectronic ions:

Electronic Configurations:

A series like this is interesting because all the ions have the same electronic structure. Notice how the ion size decreases down the series corresponding with an increase in nuclear charge. As this increases the electrons are drawn in towards the nucleus and the ion gets smaller.

$\text{N}^{3-}$	$1s^2, 2s^2, 2p^6$
$\text{O}^{2-}$	$1s^2, 2s^2, 2p^6$
$\text{F}^-$	$1s^2, 2s^2, 2p^6$
$\text{Na}^+$	$1s^2, 2s^2, 2p^6$
$\text{Mg}^{2+}$	$1s^2, 2s^2, 2p^6$
$\text{Al}^{3+}$	$1s^2, 2s^2, 2p^6$

High charge density - highly polarising

There are other isoelectronic series. Have a look at,  $\text{P}^{3-}$ ,  $\text{S}^{2-}$ ,  $\text{Cl}^-$ ,  $\text{K}^+$ ,  $\text{Ca}^{2+}$ ,  $\text{Sc}^{3+}$   
212pm 81pm  
2, 8, 8

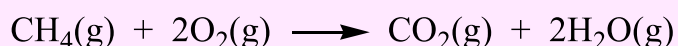
# Theme Five

## Oxidation & Reduction

In 1771-2 the Swedish chemist, **Karl Scheele**, prepared oxygen (which he referred to as *fire-air*) by heating various substances which readily undergo thermal decomposition.

Mercury(II) oxide was one such substance. In 1774 the English chemist, **Joseph Priestley**, working independently, also prepared the gas by heating the same mercury oxide. He named it *dephlogisticated* air. Priestley is usually credited with discovering oxygen because Scheele's results were not published until 1777! The French chemist **Antoine Lavoisier** prepared and investigated the properties of fire-air or dephlogisticated air and renamed it **oxygen** (from the Greek meaning, *acid*

*former*). He gave it this name because he believed it to be a constituent of all acids. We now know that not all acids contain oxygen. Chemical reactions in which oxygen combines with an element or a compound are referred to **oxidation** reactions. A good example is the combustion of methane (the main constituent of natural gas).



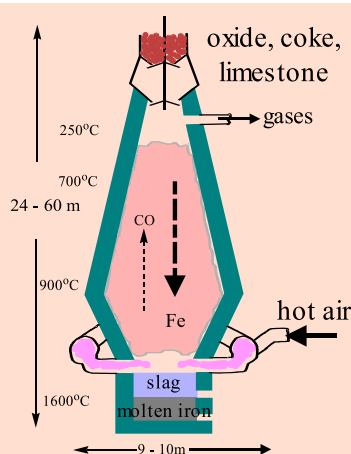
Methane is oxidised. Carbon dioxide and water are the oxidation products. We say the oxygen has been reduced. Oxidation and reduction go hand in hand; they are complementary. You can't have one without the other!

Since Lavoisier's time, the definition of oxidation has broadened. In addition to reactions in which oxygen is added (or hydrogen removed), reactions involving removal of electrons and reactions involving an increase in *oxidation number* are also considered to be oxidation reactions.

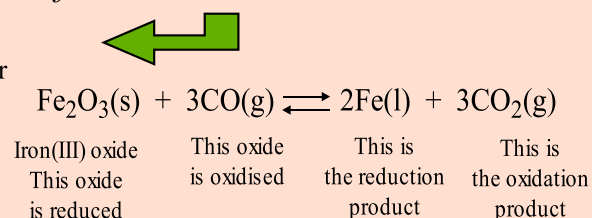
The burning of wood, coal and oil are redox reactions. The fuel is oxidised and the oxygen reduced.

The corrosion of metals like iron, in air, are redox processes. The metals are oxidised.

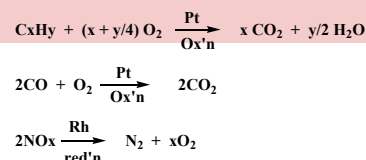
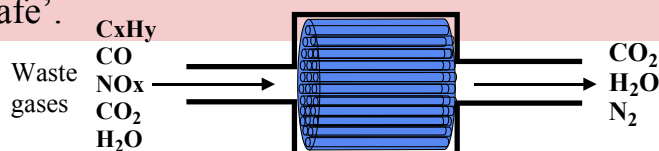
A good example of an important redox process is the conversion



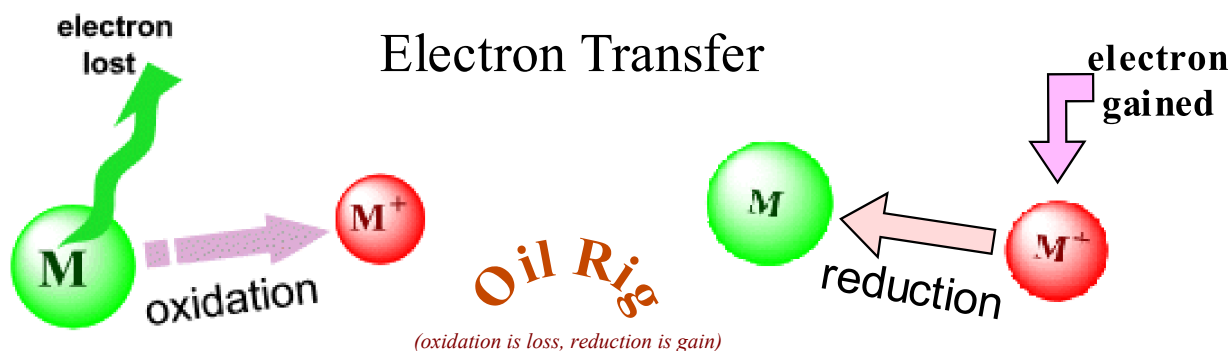
of metal oxides, like haematite (iron(III) oxide), to the metal. In the case of iron, the process may be carried out in a *blast furnace*.



A good, modern, example of the application of oxidation and reduction is in the **catalytic converter** as used in cars and commercial vehicles. The gases from the combustion cylinders pass into the exhaust system which incorporates a chamber (ie, the catalytic converter) containing a ceramic block of narrow tubes. These tubes are coated with aluminium oxide impregnated with platinum and rhodium catalysts. At the temperature of the exhaust system the waste gases are rendered 'environmentally safe'.



## Oxidation & Reduction

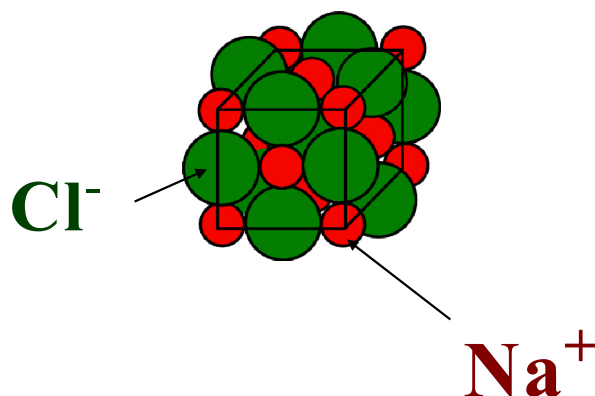


In this reaction,  
sodium atoms lose  
electrons to chlorine  
atoms.

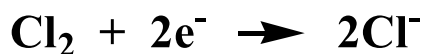
The sodium is  
oxidised and the  
chlorine is reduced.

The chlorine is the  
**oxidising agent**

Sodium chloride is a  
giant aggregate of sodium  
and chloride ions.

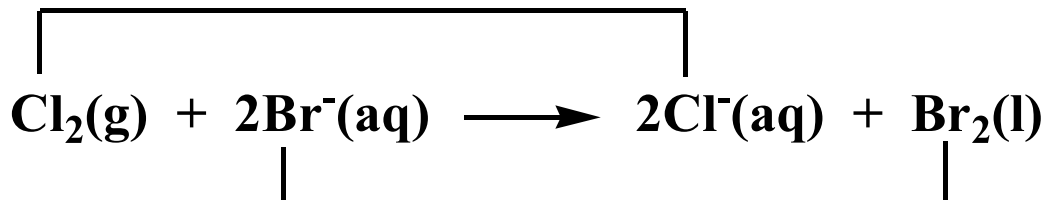


A further example:-

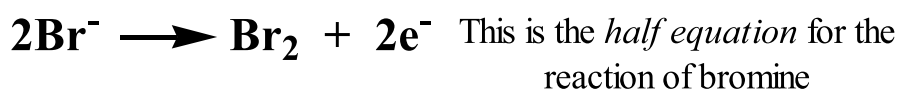


**electron gain - reduction**

This is the *half equation* for the  
reaction of chlorine



**electron loss - oxidation**



Overall ionic equation.  
Can be obtained by adding  
together the two half equations



## Oxidation & Reduction

# Oxidation Number

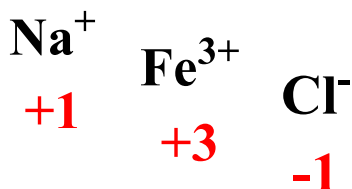
W.M. Latimer (1938)

This was introduced to rationalize oxidation & reduction definitions and to help decide whether reactions involving molecular substances, as well as ionic substance, can be classified as redox.

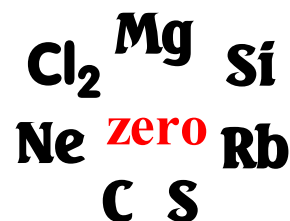
In terms of oxidation number,  
oxidation is when the oxidation number of an atom **increases** (goes more positive),  
reduction is when the oxidation number **decreases** (goes more negative).

In order to use this definition a set of **rules** need to be applied to decide the *value* of the oxidation numbers of atoms involved in chemical reactions.

Monatomic ions - charge on the ion



Elements - zero



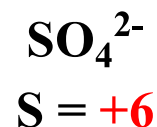
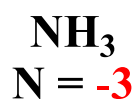
Hydrogen in its compounds:  
**+1**

Except in NaH in which its **-1**

Oxygen in its compounds:  
**-2**

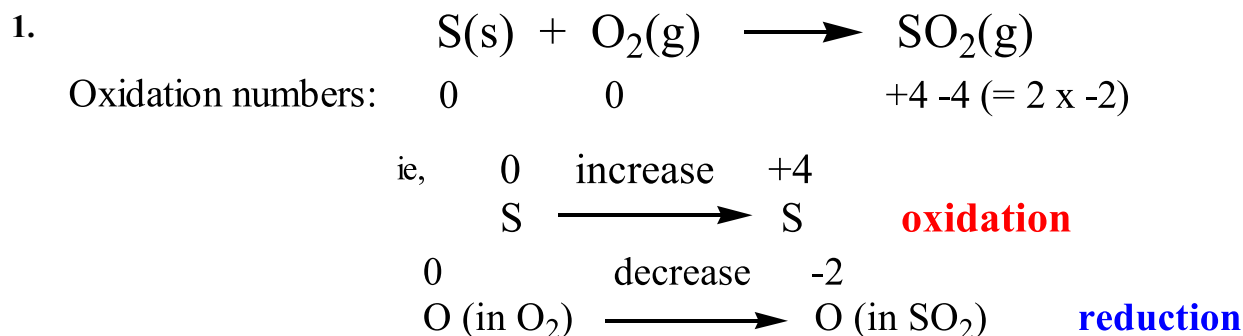
Except in  $\text{H}_2\text{O}_2$   
where its **-1**  
and  $\text{F}_2\text{O}$  where its **+2**

The algebraic sum of the individual oxidation numbers is zero in molecules and the charge on the ion, in ions.

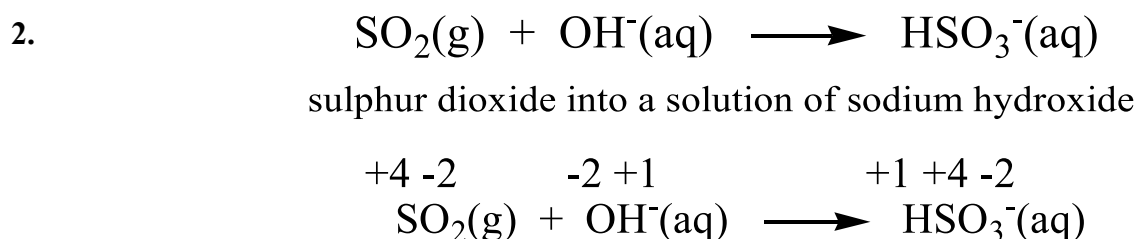


## Oxidation & Reduction

### Applying oxidation number to reactions



There has been a change in oxidation number. This is, therefore, a redox reaction.



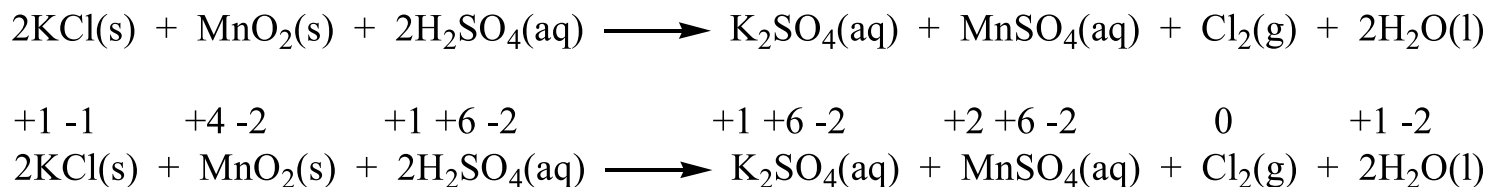
The atoms of the elements have not changed their oxidation numbers.

**This is not a redox reaction.**

3. When a metal chloride is heated with manganese dioxide and concentrated sulphuric acid chlorine is evolved. This reaction has been used as a test for chlorides.

Is this type of reaction redox?

eg,



Cl      increase, therefore **oxidised**

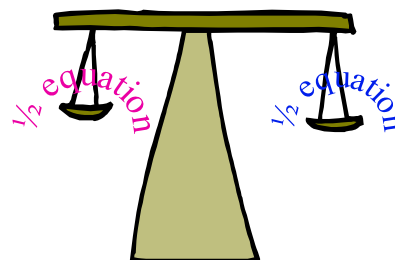
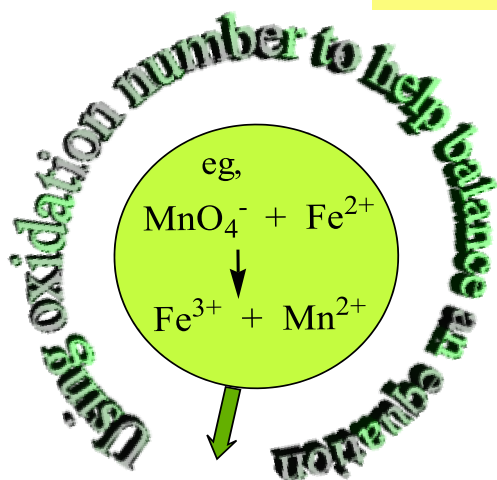
Mn      decrease, therefore **reduced**

Other elements, no change in oxidation number

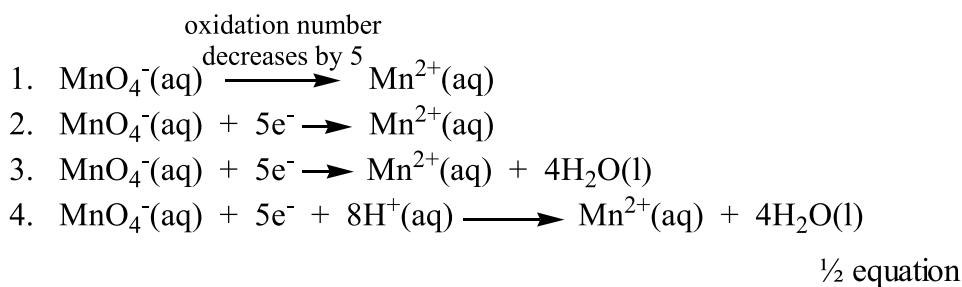
**The reaction is a redox reaction**



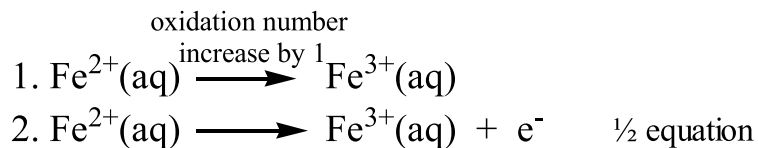
## Oxidation & Reduction



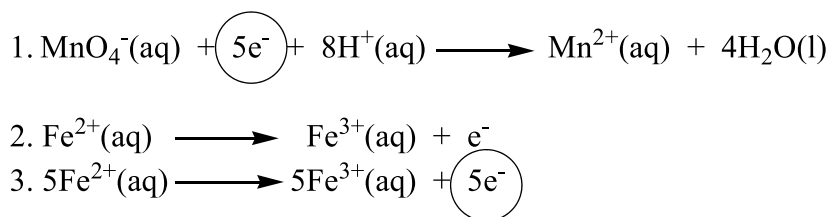
A. Obtain balanced half equation for permanganate



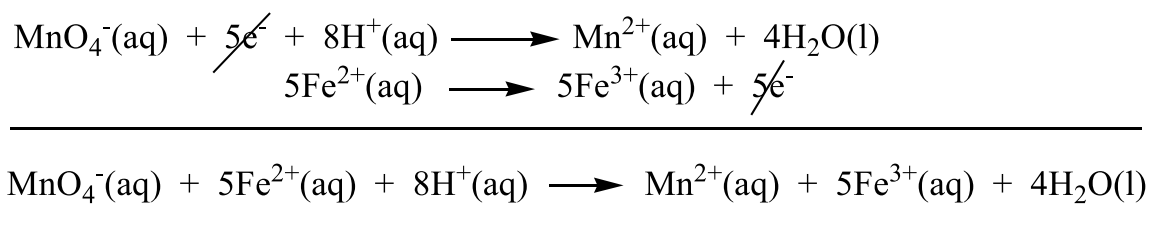
B. Obtain the balanced half equation for iron(II)



C. Compare the two half equations and balance with respect to changes in oxidation number.



D. Add the two half equations cancelling out the electrons



Final equation

# Theme Six

## Halogens

### The Halogens (Group 7)

The name, halogen, derives from the Greek meaning *salt-producer*. Many of their compounds are found as solutes in sea water.

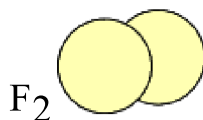
The halogens are *diatomic*.

The halogens are toxic.



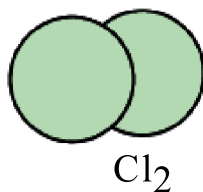
#### fluorine

French chemist  
H. Moissan (1886)



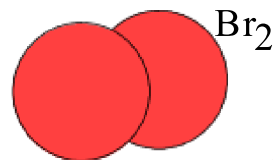
#### chlorine

Swedish chemist  
C.W. Scheele (1774)



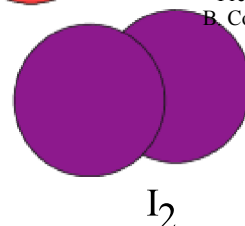
#### bromine

German chemist C.J. Lowig (1825)  
& French chemist A. Balard (1826)



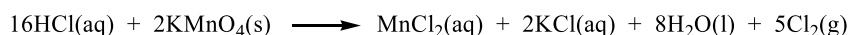
#### iodine

French chemist  
B. Courtois (1811)

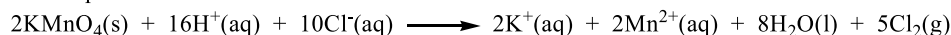


**Chlorine** is typical of the group. It is manufactured by the electrolysis of aqueous sodium chloride using either the *membrane* (diaphragm) *cell* or the *mercury cathode cell*.

In the laboratory it is made by oxidising concentrated hydrochloric acid with potassium permanganate (refer to the practical work pages).



Ionic equation:



Halogens					
Name	Symbol	Mp <sup>0</sup> C	Bp <sup>0</sup> C	Physical	Colour
fluorine	F	-220	-188	gas	pale yellow
chlorine	Cl	-102	-35	gas	pale green
bromine	Br	-7	59	volatile liquid	reddish-
iodine	I	113	184	solid	brown-black

#### *Like dissolves like.*

The halogens are very soluble in the organic solvent tetrachloromethane. The solutions acquire the colour of the gaseous elements.

**When heated under normal conditions, iodine sublimes to give a violet vapour**



## Halogens

The halogens have similar chemical properties because they have similar outer electronic structures.



Halogen	Electron configuration	Atomic radius (pm)	1st ionisation energy (kJ mol <sup>-1</sup> )	1st electron affinity (kJ mol <sup>-1</sup> )
fluorine	2,7	72	1680	-348
chlorine	2,8,7	99	1260	-364
bromine	2,8,18,7	114	1140	-342
iodine	2,8,18,18,7	133	1010	-314

Increase in atom size  
down the group.

Decrease in ionisation  
energy down the  
group.

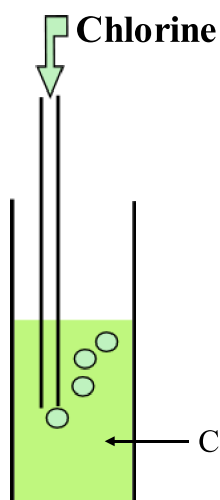
Decrease in electron  
affinity down the  
group.

These are electronegative elements.

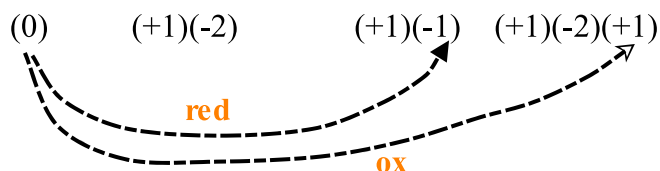
They have a strong affinity for electrons and readily form uni-negatively charged ions.

However, these properties and reactivity decrease down the group as atom size increases.

## With water



Chlorine dissolves in water (about 7g per litre @ rtp). Some of the chlorine (about a third) reacts with the water producing hydrochloric and chloric(I) acids. The reaction is reversible and it provides an example of a **disproportionation** reaction.



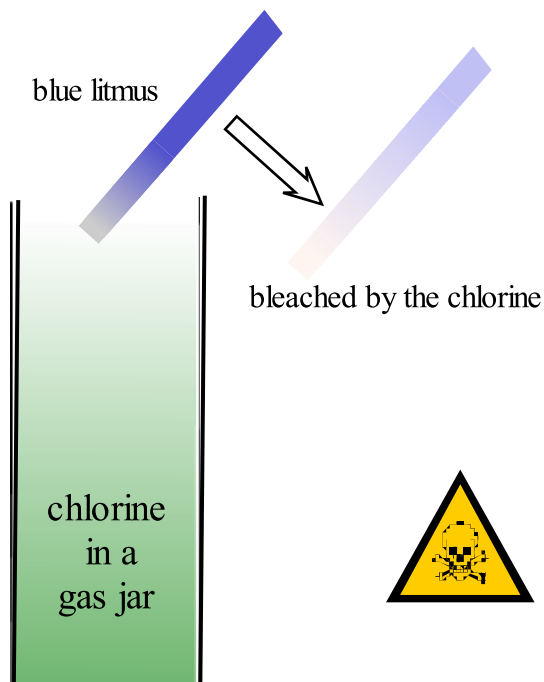
Chlorine undergoes both oxidation & reduction

Bromine is more soluble in water than chlorine (about 34g per litre @ rtp). Iodine is the least soluble (about 0.35g per litre). However, bromine and iodine **react** only to a small extent with water.

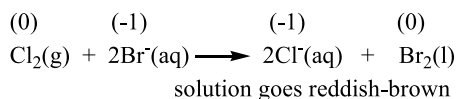
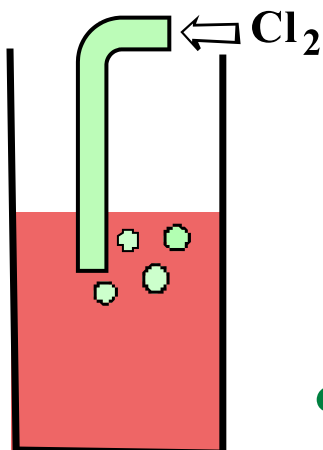
# Halogens

## Tests for chlorine:

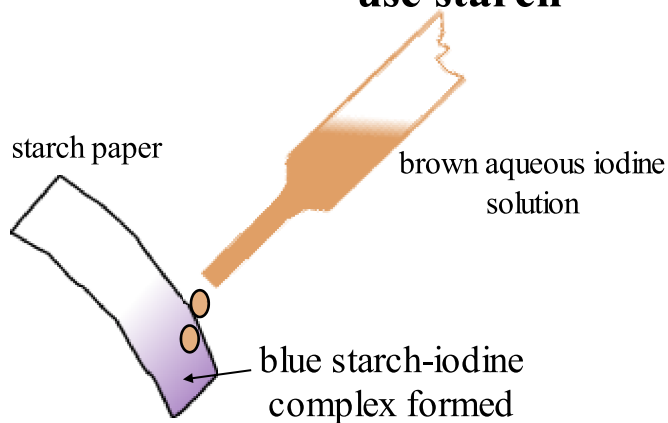
1. blue litmus



2.



## Test for Iodine - use starch



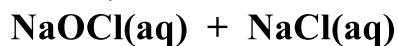
Chlorine gas bubbled into colourless potassium bromide solution  
- - solution turns red-brown.

This is an example of a **displacement reaction**.

(Bromine is obtained from sea water (0.007% bromide) by this process)



↓ cold



**bleach**

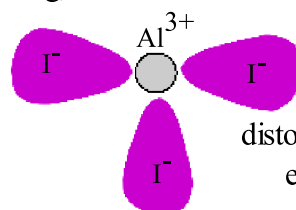


Oxidising Power:  $\text{F} > \text{Cl} > \text{Br} > \text{I}$

Electronegativity:  $\text{F} > \text{Cl} > \text{Br} > \text{I}$

Polarisability:  $\text{I}^- > \text{Br}^- > \text{Cl}^- > \text{F}^-$

eg,

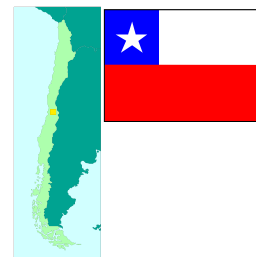


distortion of the iodide electron clouds

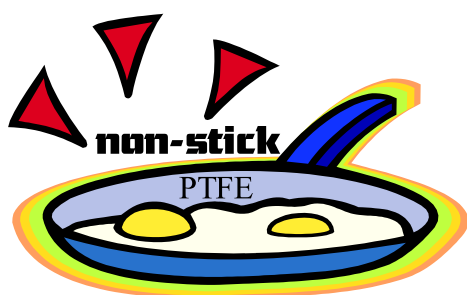
## Halogens



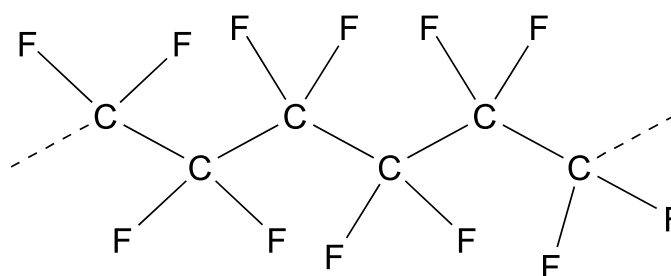
As late as 1940, some iodine was still obtained from seaweed. Various kelps were dried and burned and the iodine extracted from the iodides in the residual ash. Today, iodine is extracted from sodium iodate which is found in Chile saltpetre (a mixture of sodium iodate and nitrate (1:500))



Silver bromide is important in photography. It is used in the development of the image on the photographic film.



An important use of fluorine is in the manufacture of Teflon. This is a high melting plastic which has non-stick properties. It is made from polytetrafluoroethene

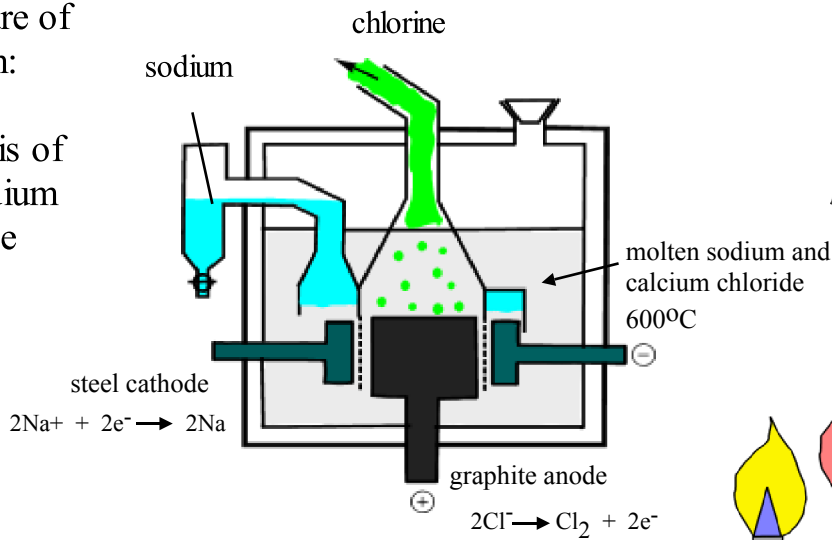


## The s-Block Elements

I	II	The Periodic Table																III	IV	V	VI	VII	0
1 H																							2 He
3 Li	4 Be																	5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg																	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr						
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe						
55 Cs	56 Ba	57-71	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn						
87 Fr	88 Ra	89-																p-block					
s-block																							

## Manufacture of sodium:

## Electrolysis of fused sodium chloride



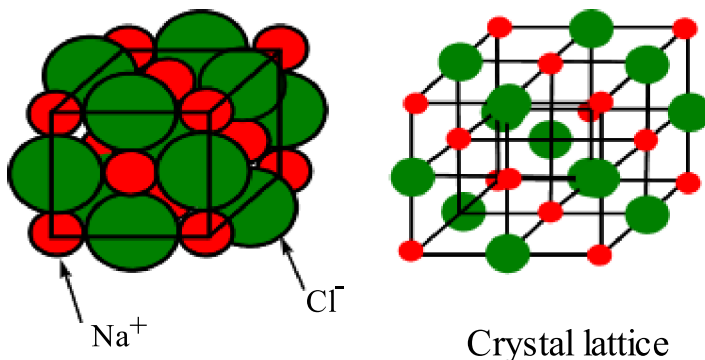
## Flame Test - which metal gives which colour?

Label the  
diagram!

Plaster of,  
.....?



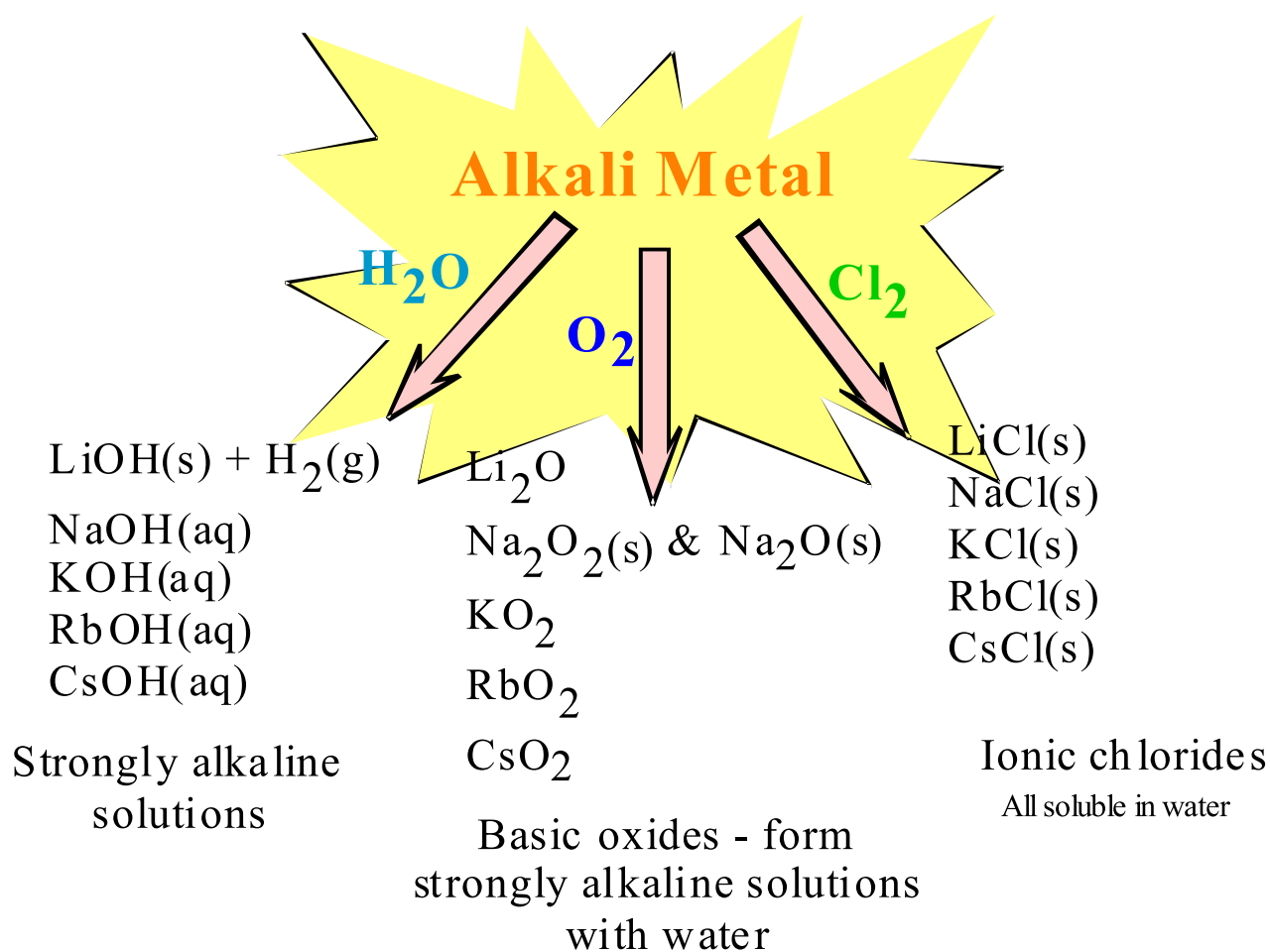
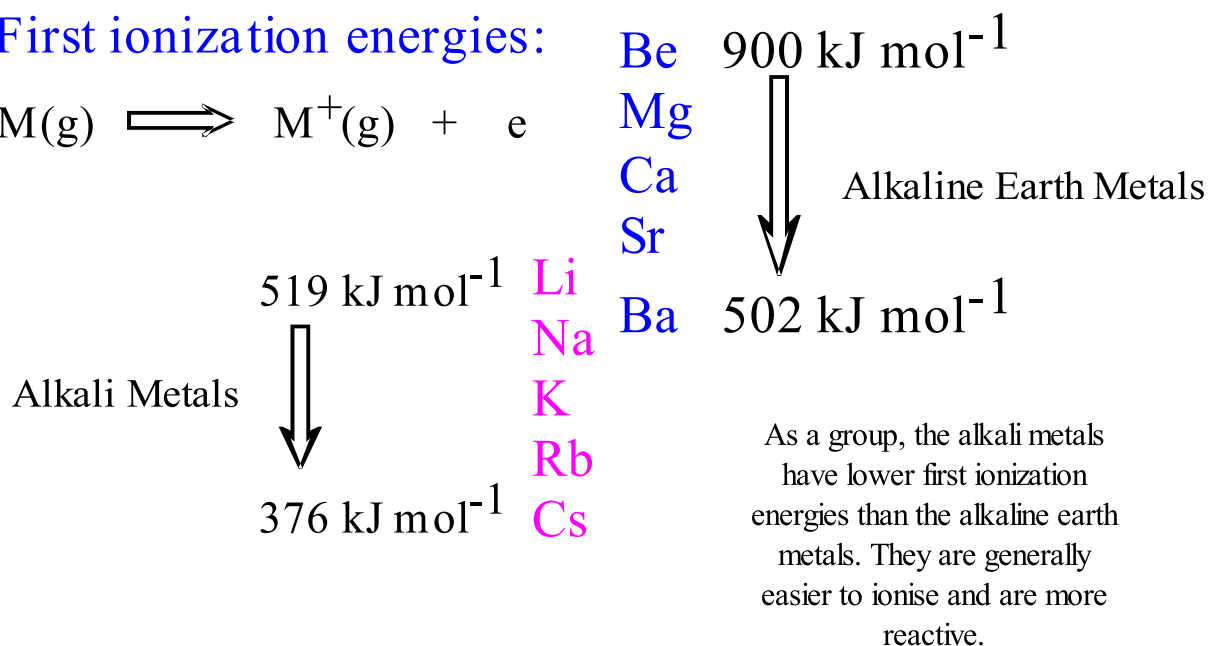
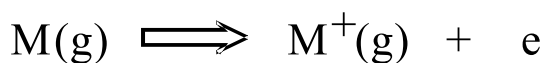
## Sodium chloride crystal lattice



## Crystal lattice

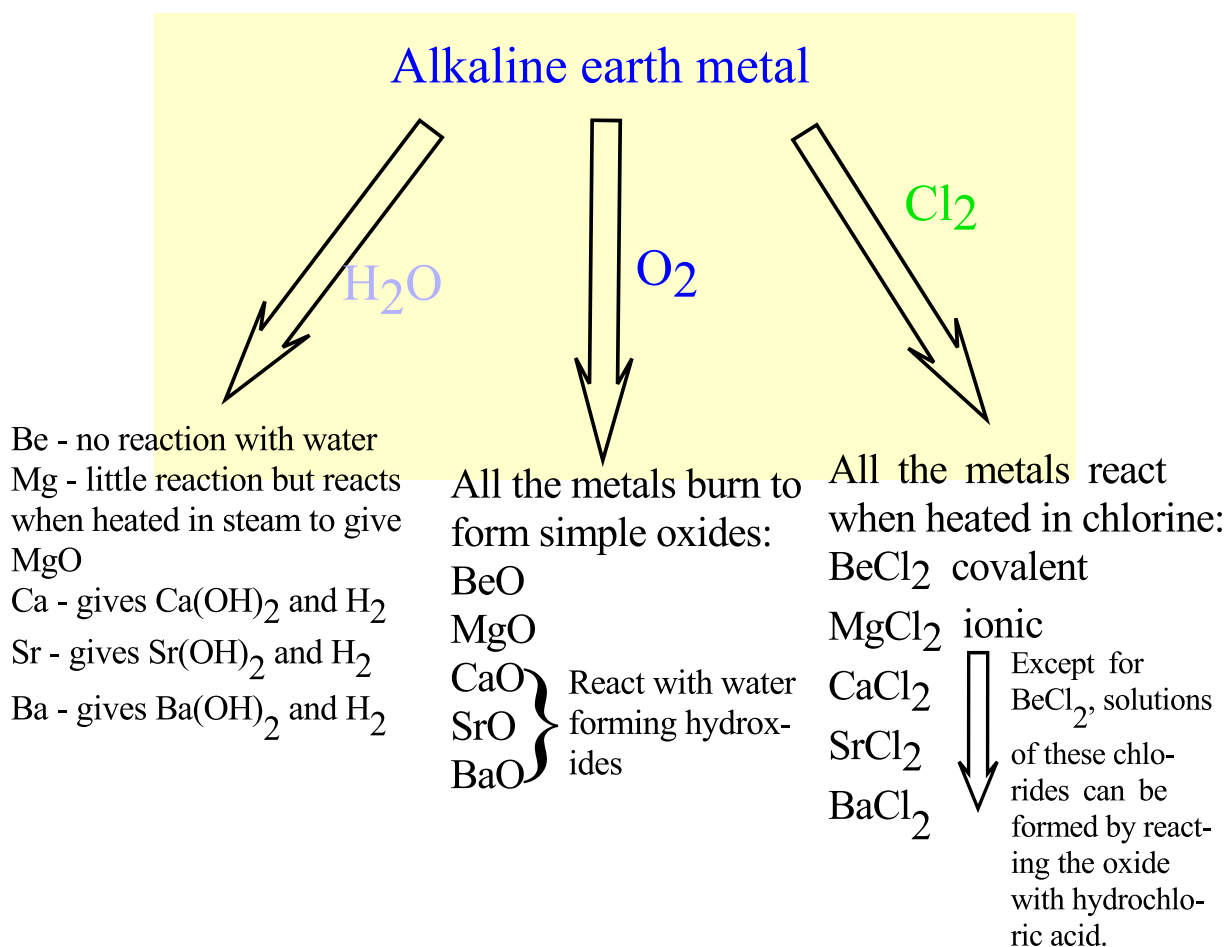
## The s-Block Elements

First ionization energies:



(lilac = potassium, green = barium,  
red (brick) = calcium, yellow = sodium)

## The s-Block Elements

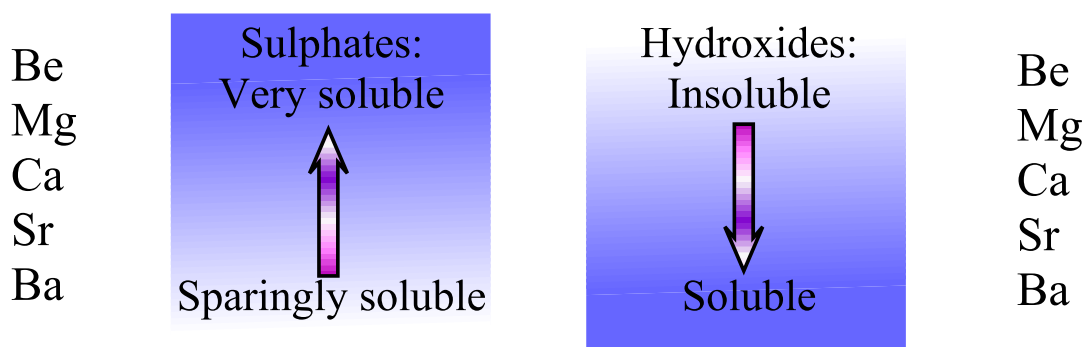


## OXIDATION NUMBERS

Alkali metals: 0 and +1

Alkaline earth metals: 0 and +2

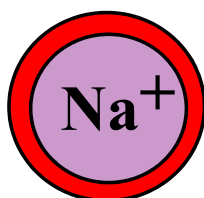
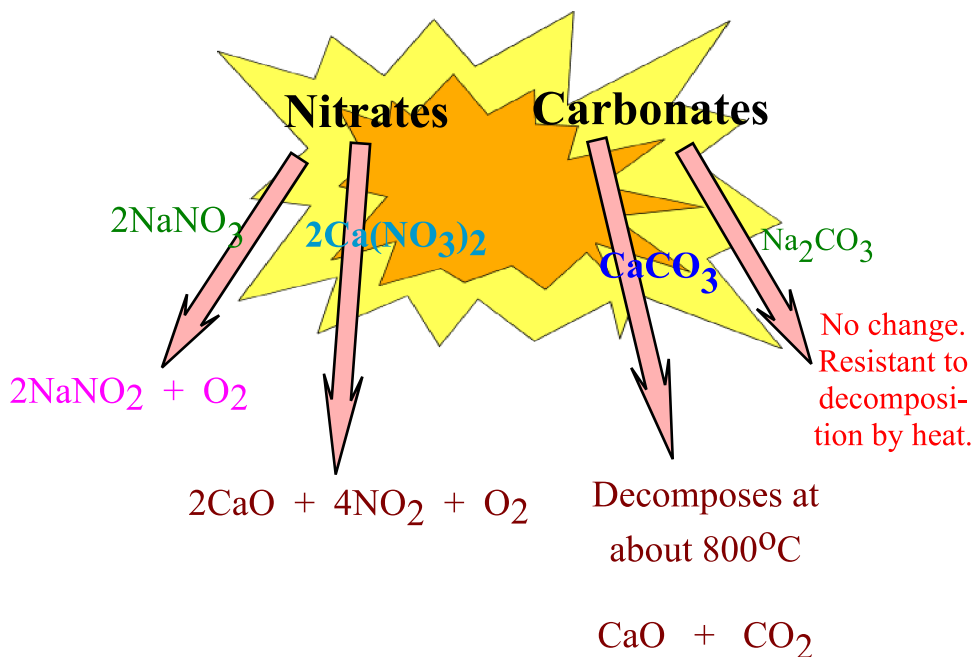
## Solubility of the sulphates and hydroxides of Group 2:



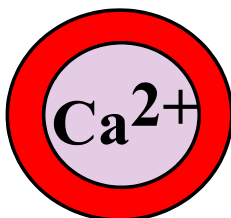


## The s-Block Elements

### Thermal Decomposition



Weakly polarising



More polarising

Highly charged cations deform (polarise) anions and encourage their decomposition

*This is why the alkaline earth metal salts decompose on heating more readily than those of the alkali metals.*

The s-block elements are good conductors of heat and electricity.  
Sodium is used as a heat exchanger in some nuclear reactors.  
Magnesium is alloyed with aluminium and used in aircraft construction.

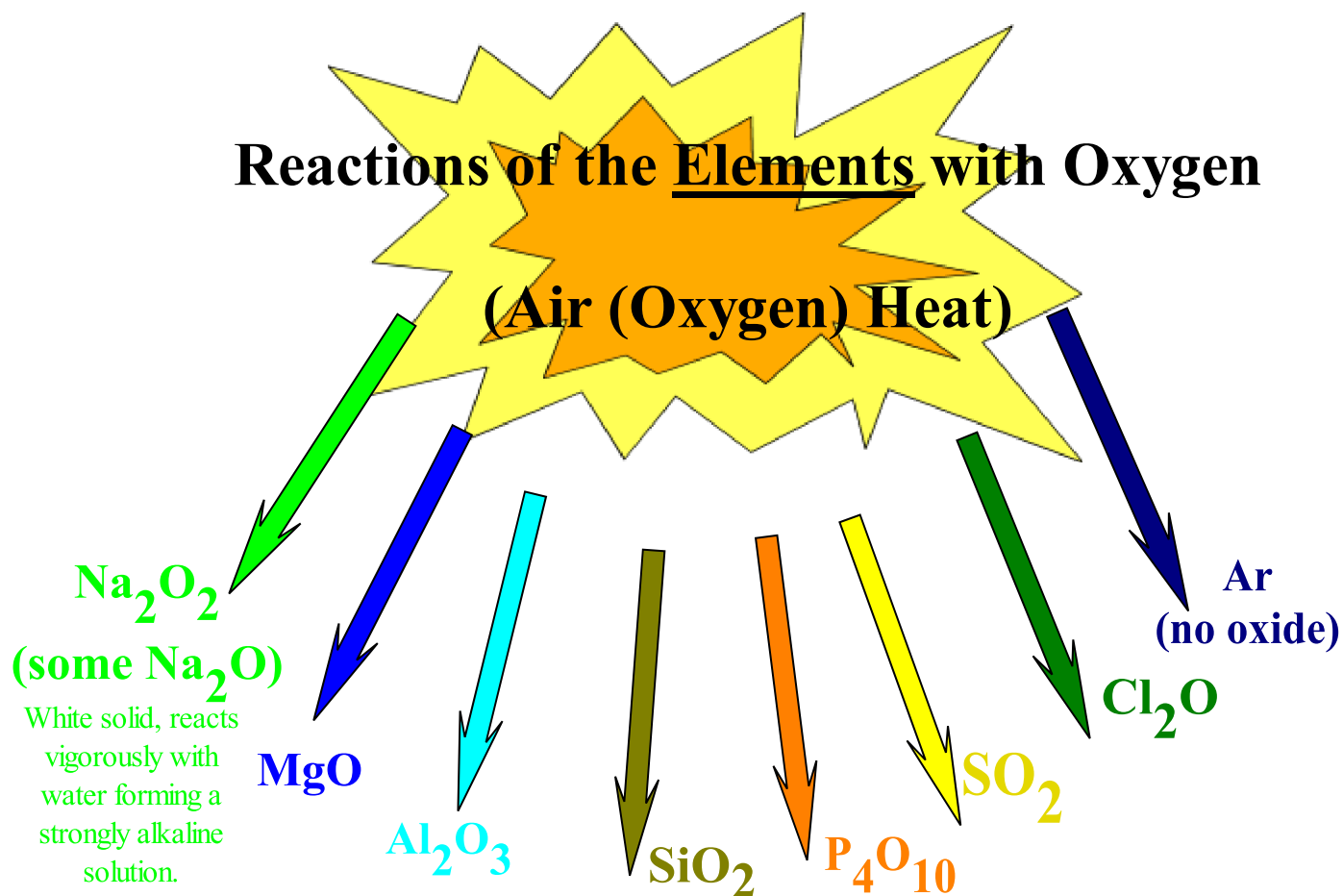


## The Period 3 Elements

Element	Na	Mg	Al	Si	P	S	Cl	Ar
Electron Configuration	2,8,1	2,8,2	2,8,3	2,8,4	2,8,5	2,8,6	2,8,7	2,8,8
Outer Shell	3s <sup>1</sup>	3s <sup>2</sup>	3s <sup>2</sup> ,3p <sup>1</sup>	3s <sup>2</sup> ,3p <sup>2</sup>	3s <sup>2</sup> ,3p <sup>3</sup>	3s <sup>2</sup> ,3p <sup>4</sup>	3s <sup>2</sup> ,3p <sup>5</sup>	3s <sup>2</sup> ,3p <sup>6</sup>
Electronegativity	0.9	1.2	1.5	1.8	2.1	2.5	3	-
1st Ionisation energy (kJ mol <sup>-1</sup> )	494	736	577	786	1060	1000	1260	1520
Atom radius (pm)	186	160	143	117	110	104	99	154
Ion radius (pm)	95	65	50	-	-	184	181	-

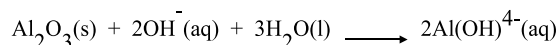
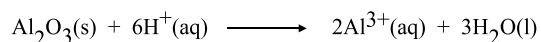
Note how these properties change going from Na to Ar across the period.  
There is an overall decrease in atom size as the third shell is filled. It gets progressively more difficult to ionise the atoms.

<div style="display: flex; align-items: center; justify-content: space-between;"> <span>←</span> <span>Across the Period</span> <span>→</span> </div>		
Metals, low electronegativity, metallic structures	Non-metals, medium electronegativity, covalent lattices	Non-metals, high electronegativity, molecular structure
Na, Mg, Al	Si, P(P <sub>4</sub> ), S(S <sub>8</sub> )	Cl(Cl <sub>2</sub> ), Ar
M.p °C: 98, 650, 660	M.p °C: 1410, 44, 119	M.p °C: -101, -189



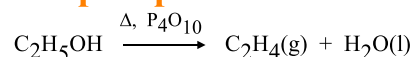
**MgO:** White solid, giant aggregate of metal ions and oxide ions. A basic oxide: reacts with mineral acid forming a salt and water. Reacts feebly with water forming a v slightly alkaline solution. Magnesium hydroxide is sparingly soluble in water.

**Al<sub>2</sub>O<sub>3</sub>:** Forms a tough protective layer on the metal. It is an amphoteric oxide. When freshly made it reacts with both acids and alkalis.



**SiO<sub>2</sub>:** White solid. Acidic oxide. Forms a soluble silicate when heated with a concentrated solution of sodium hydroxide.

**P<sub>4</sub>O<sub>10</sub>:** This is a low melting white solid. Acidic oxide. Gives phosphoric acid with water. Used as a drying agent & dehydrating agent.



**SO<sub>2</sub>:** This is a colourless toxic gas. Acidic oxide. Dissolves in water forming sulphuric (IV) acid. Oxidised to sulphur trioxide which is used to manufacture sulphuric(VI) acid.

**Cl<sub>2</sub>O:** This is a yellow/brown poisonous gas. Acidic oxide. It dissolves readily in water forming a solution containing chloric(I) acid (HOCl(aq)).

# Transition Metals

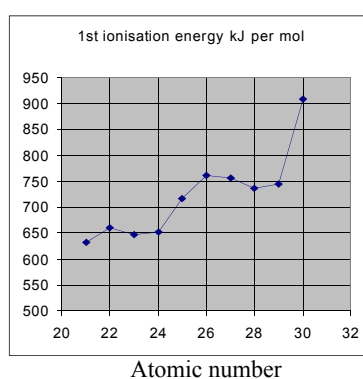
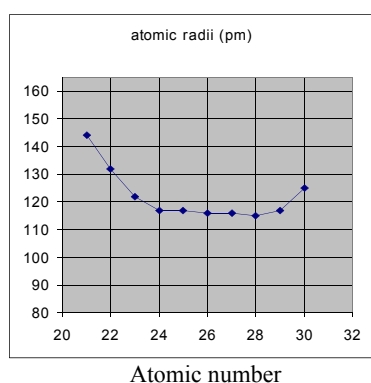
																1.01 H 1	4.00 He 2
6.94 Li 3	9.01 Be 4											10.81 B 5	12.01 C 6	14.01 N 7	16.00 O 8	19.00 F 9	20.18 Ne 10
22.99 Na 11	24.31 Mg 12											26.98 Al 13	28.09 Si 14	30.97 P 15	32.07 S 16	35.45 Cl 17	39.95 Ar 18
39.10 K 19	40.08 Ca 20	44.96 Sc 21	47.87 Ti 22	50.94 V 23	52.00 Cr 24	54.94 Mn 25	55.85 Fe 26	58.93 Co 27	58.69 Ni 28	63.55 Cu 29	65.39 Zn 30	69.72 Ga 31	72.64 Ge 32	74.92 As 33	78.96 Se 34	79.90 Br 35	83.80 Kr 36
85.47 Rb 37	87.62 Sr 38	88.91 Y 39	91.22 Zr 40	92.91 Nb 41	95.94 Mo 42	99 Tc 43	101.07 Ru 44	102.91 Rh 45	106.42 Pd 46	107.87 Ag 47	112.41 Cd 48	114.82 In 49	118.70 Sn 50	121.75 Sb 51	127.60 Te 52	126.90 I 53	131.29 Xe 54
132.91 Cs 55	137.33 Ba 56	138.91 La 57 *	178.49 Hf 72	180.95 Ta 73	183.84 W 74	186.21 Re 75	190.23 Os 76	192.22 Ir 77	195.09 Pt 78	196.97 Au 79	200.59 Hg 80	204.38 Tl 81	207.2 Pb 82	208.98 Bi 83	209 Po 84	210 At 85	222 Rn 86
223 Fr 87	226.03 Ra 88	227.03 Ac 89 **	261 Rf 104	262 Db 105	266 Sg 106	264 Bh 107	277 Hs 108	268 Mt 109	271 Ds 110	272 Rg 111							
* Lanthanides		140.12 Ce 58	140.91 Pr 59	144.24 Nd 60	147 Pm 61	150.36 Sm 62	151.69 Eu 63	157.25 Gd 64	158.93 Tb 65	165.50 Dy 66	164.93 Ho 67	167.26 Er 68	168.93 Tm 69	173.04 Yb 70	174.97 Lu 71		
** Actinides		232.04 Th 90	231.04 Pa 91	238.03 U 92	237.05 Np 93	244 Pu 94	243 Am 95	247 Cm 96	247 Bk 97	251 Cf 98	252 Es 99	257 Fm 100	257 Md 101	259 No 102	261 Lr 103		

First transition series

**Transition metals** are defined as elements which have at least one of their ions with a partly filled d-sublevel. In the first transition series all of the elements satisfy this definition *except* for scandium and zinc. These metals only exhibit one oxidation state each (ie, Sc(III) and Zn(II)). In Sc(III) there are no electrons in the 3d level. In Zn(II) the 3d level is full.

Since both the 3d and 4s electrons are available for bonding purposes the metals, Ti to Cu, exist in a large number of oxidation states. Manganese, for example, shows seven oxidation states, II and VII being the most common.

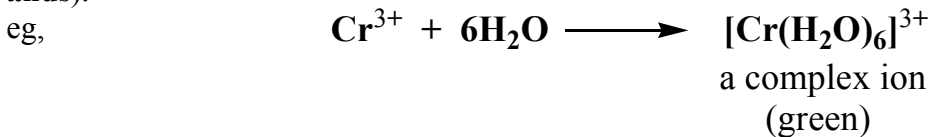
Atomic radii and ionisation energies change only gradually across the series



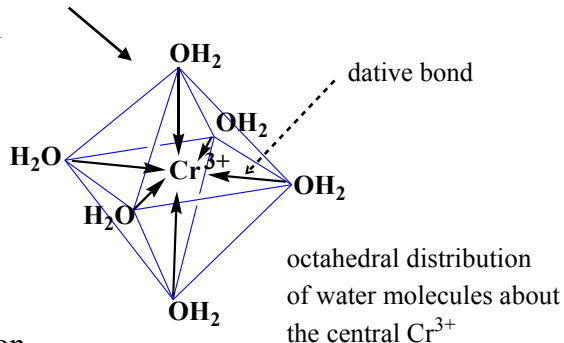
ture of 4190°C and titanium has a low density of 4.49 g cm<sup>-3</sup>

## Complex ions

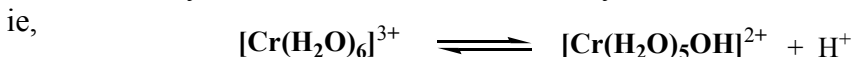
Ions in which the transition metal atom is bonded to a number of electron donating atoms or groups (known as **ligands**).



The number of ligands attached to the central metal ion is known as the *co-ordination number*. In this example the co-ordination number is six.

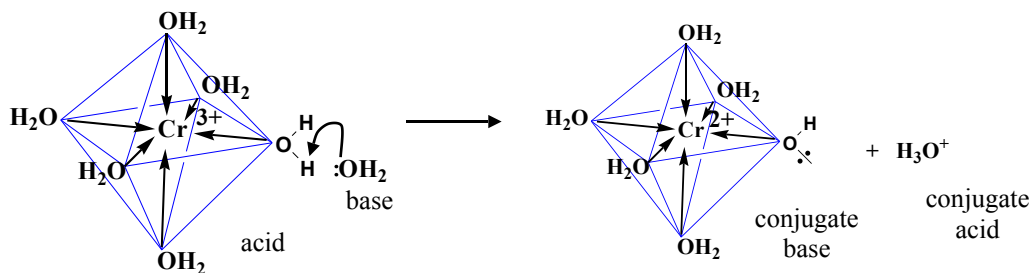


Solutions of hydrated ions, like this, are weakly acidic due to dissociation.

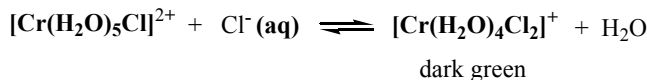
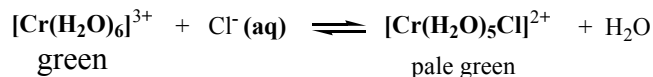


Generally, the higher the charge on the metal ion the more acidic the solution. The reaction is an example of a *deprotonation reaction*.

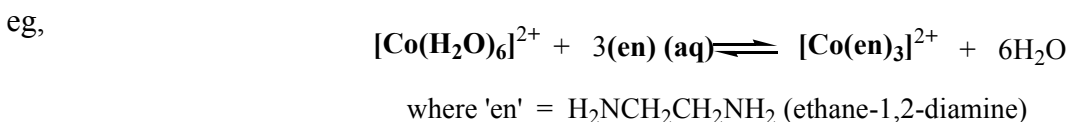
(note: the blue lines are **not** bonds - they are included simply to illustrate & emphasise the octahedral shape!)



Water molecules are weakly held ligands which can be replaced by stronger ligands. Adding brine ( $\text{NaCl(aq)}$ ), for example, to a solution of the hydrated  $\text{Cr}^{3+}$  ion produces a colour change as chloride ions replace the water molecules.

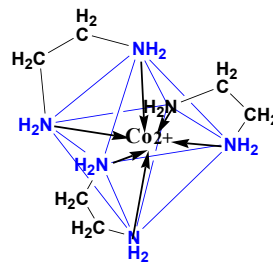


Some ligands can occupy more than one co-ordinating position. They are called **chelating agents** (Greek - *crab-like*). Oxalate (ethanedioate) ion and ethane-1,2-diamine are good examples. These are called *bidentate* ligands because they form two dative links to the central metal ion.

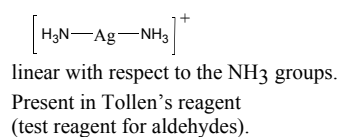
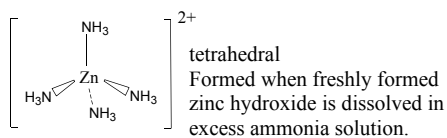
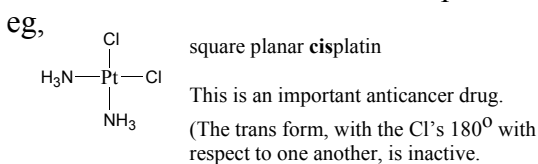


Just as a matter of interest, the substance EDTA (ethylenediaminetetraacetic acid) forms **six** dative links (hexadentate) to transition metal ions. It is an important reagent in the analysis of transition metals & in the treatment of lead poisoning.

Haem, a component of blood haemoglobin, is a chelate of iron(III). Chlorophyll is a chelate of magnesium.



**Note:** not all transition metal complexes have octahedral structures.

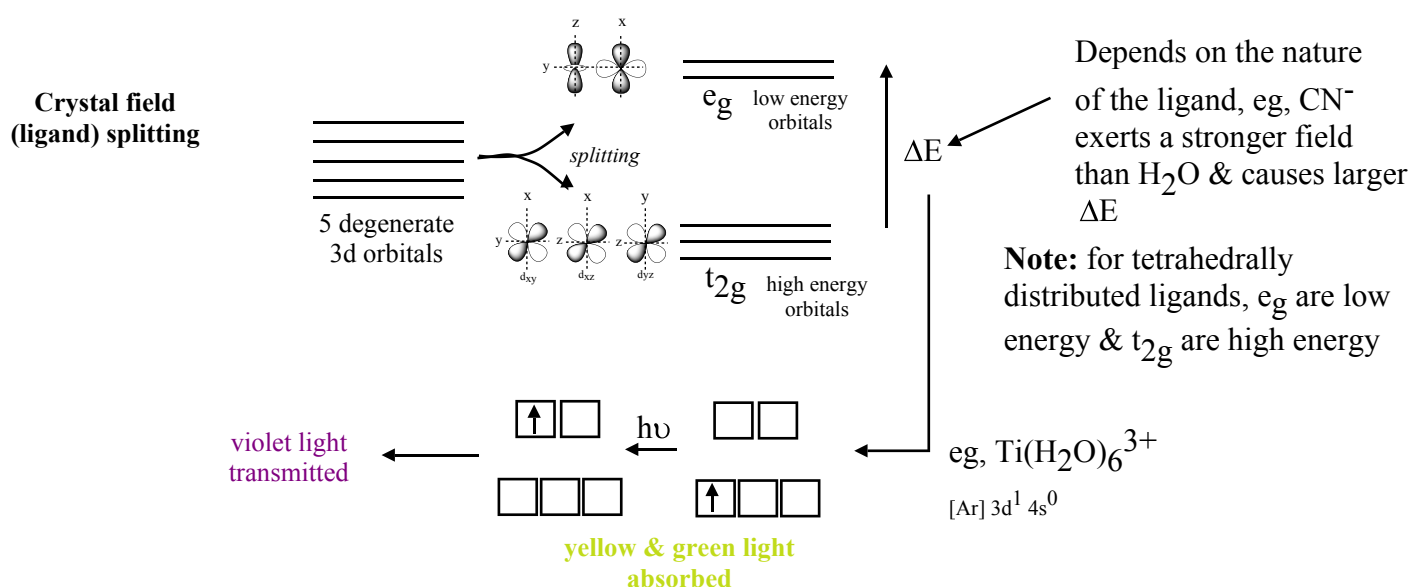


# Coloured Compounds

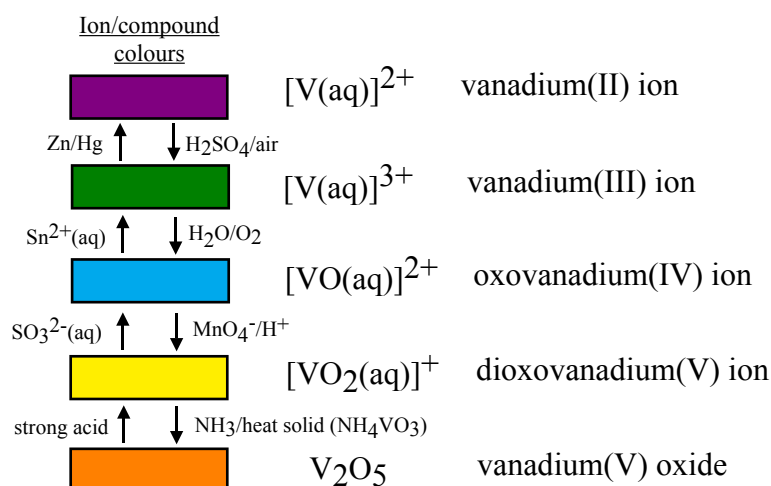
Another characteristic feature of transition metals is that they form a large number of coloured compounds.  
eg,

$\text{CuSO}_4$ (hydrated)	-----	blue
$\text{Ni}(\text{NO}_3)_2$	"	----- green
$\text{CrCl}_2$	"	----- blue
$\text{MnCl}_2$	"	----- pink
$\text{FeSO}_4$	"	----- pale green
$\text{TiCl}_3$	"	----- violet

The colour is due to the metal ion absorbing visible light, eg,  $\text{Cu}^{2+}(\text{aq})$  absorbs red/yellow light leaving residual blue light. The light absorbed has just the right amount of energy to promote the d electrons to higher energy levels. If the metal ion is surrounded by six *octahedrally* positioned ligands then the d orbitals separate (or split) into two sets, high energy ( $e_g$ ) and low energy ( $t_{2g}$ ). The high energy orbitals are those closest to the ligands:



## Colourful Vanadium Chemistry (summary):



# Catalysts

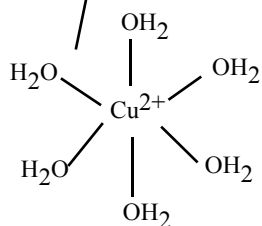
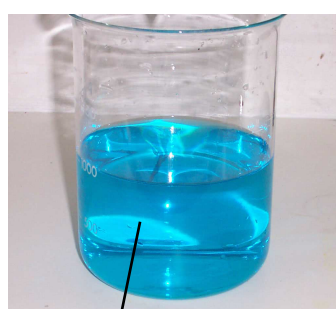
These are substances which alter the rate of a reaction by providing an alternative reaction pathway. Positive catalysts provide a less energetic route. Transition metals and their compounds provide very useful industrial catalysts.

eg,	$V_2O_5$	Contact process
	$Fe/Fe_2O_3$	Haber process
	Ni	Reduction of alkenes
	Pt/Rh	Ostwald process (oxidation of ammonia)
	$TiCl_3$	Polymerisation of alkenes

These are examples of heterogeneous catalysts. They are in a different phase to the reactants. This situation is generally preferred for industrial processes because, (i) they can be easily separated from the products, (ii) they are generally more suited to continuous industrial processes (rather than batch) and (iii) they can show stereospecificity (ie., produce one optical isomer rather than the other!).

## Some Copper Chemistry

Copper(II) sulphate solution



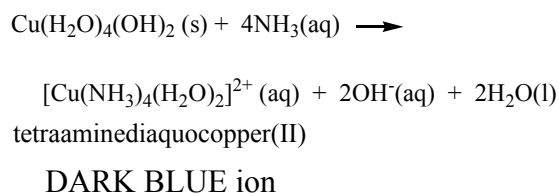
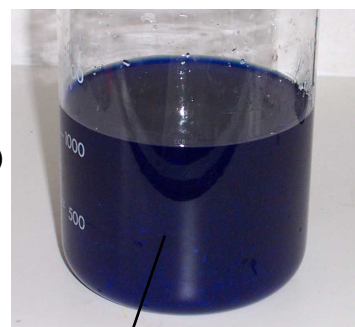
octahedral hexaaquacopper(II) ion  
BLUE

$\xrightarrow{\text{NaOH(aq)}}$



hydrated  $\text{Cu}(\text{OH})_2$   
PALE BLUE solid

$\xrightarrow{\text{NH}_3(\text{aq})}$



## Manganese(II) and cobalt(II) solutions:

Both are pink but they can be distinguished using sodium hydroxide solution & ammonia solution:

Adding sodium hydroxide solution to a solution of manganese(II) ion produces light brown manganese(II) hydroxide. This is insoluble in excess  $\text{NaOH}(\text{aq})$  and insoluble in ammonia solution.

With sodium hydroxide solution, cobalt(II) solutions give a blue precipitate of cobalt(II) hydroxide.

This turns pink on standing. The hydroxide is soluble in ammonia solution forming the blue  $[\text{Co}(\text{NH}_3)_6]^{2+}$  ion

Which is readily oxidised to  $[\text{Co}(\text{NH}_3)_6]^{3+}$  in air; this is yellow.

On a point of interest, hydrated cobalt(II) chloride is pink due to the octahedral  $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$  ion. On heating this ion loses some water and forms the blue tetrahedral  $[\text{Co}(\text{H}_2\text{O})_4]^{2+}$  ion. The change is reversible.

## Chromium(III) and Nickel(II) solutions:

Both are green but they can be distinguished using sodium hydroxide solution & ammonia solution:

Adding sodium hydroxide solution to chromium(III) solutions gives a green precipitate of chromium(III) hydroxide.

It is insoluble in ammonia but soluble in excess sodium hydroxide solution forming the dark green  $[\text{Cr}(\text{OH})_6]^{3-}$  ion.

With sodium hydroxide solution nickel(II) solutions give a pale green precipitate of nickel(II) hydroxide.

This is insoluble in excess sodium hydroxide solution but soluble in ammonia solution forming the deep blue  $[\text{Ni}(\text{NH}_3)_6]^{2+}$  ion.